

AN ELEMENTARY EXPERIMENTAL CHEMISTRY

JOHN BERNARD EKELEY, A.M.



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An Elementary Experimental Chemistry

BY

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P R E F A C E

THERE are two things that the study of a science should accomplish for the student: first, the development of the powers of observation; second, a knowledge of the relations of the principles and facts that underlie that science. The first may be obtained to a greater or less degree from the study of any science by the experimental method; the latter however may often be only imperfectly acquired even by those who are able to make accurate observations, but who fail to draw the correct or logical conclusions. It is the aim of this book to aid the student in accomplishing both these things. The author has found by experience (as, indeed, what teacher has not?) that all beginners in the study of chemistry are by no means natural adepts at making the necessary observations, but that by a sufficiently prolonged course of experiments their powers in this direction are easily augmented. The difficulty comes, however, in enabling the student, after the observations have been made, to form the correct conclusions, and, finally, when both these things have been done throughout the subject, to round it off into a symmetrical whole.

To accomplish these things, the author believes that it is necessary to help the student considerably by emphasizing what must especially be marked in the experiment, and to outline the course of thought that must be followed, in order that the student may feel satisfied at the close of the experi-

ment that he has made *some* definite progress at least in the process of building up his knowledge of the subject.

In Parts I. and II., this method is followed. Part I. is made up entirely of experiments dealing with *The Preparation and Properties of the Principal Elements and Compounds*. No teaching of chemical theory whatever is attempted in this part. *Qualitative* equations are given, showing merely *what kinds of matter* have been concerned in each chemical change. Words, and not chemical symbols, are used, the conventional chemical equation being left for Part II. It is hoped that the student, when he has performed the experiments of Part I., may have a good idea of the qualitative composition of the principal compounds, and of how he may distinguish them from one another. Each experiment gives him a knowledge of one or more elements or compounds; in doing this, no compound is used that has not been studied in some previous experiment. Thus the student's knowledge is built up step by step by the inductive method, and in a logical manner, until he has made, and studied the properties of those elements and compounds of inorganic chemistry with which he is most concerned.

The author believes that the study of chemical theory is most successfully carried on after the student has prepared the elements and compounds and has studied their properties. Hence Part II. is concerned entirely with *The Laws and Theories of Chemistry*. Special emphasis is laid upon the difference between *laws* and *theories*. The laws are illustrated by experiments; and the reasoning upon which we base our belief in the respective explanatory theories, which most students find so difficult, is given in full. This gives an accurate idea of theoretical chemistry instead of the hazy conceptions that are so often obtained when the theory is scattered throughout a series of merely descriptive experi-

ments. The use of chemical symbols and the writing of reactions are now taken up in detail, and numerous examples are given to illustrate the stoichiometrical relations of the elements. The plan of deferring the writing of reactions until late in the course is in line with the most recent suggestions of the best teachers of the science. Throughout the book the revised spelling of chemical names is used. The illustrations of apparatus in Parts I. and II. will save the teacher valuable time that would otherwise be used in answering questions on manipulation.

The author wishes to acknowledge his indebtedness to Prof. J. F. McGregory of Colgate University, and to Dr. Albert C. Hale of the Boys' High School, Brooklyn, for many valuable suggestions, and to his colleague at St. Paul's, Mr. Arthur De L. Ayrault, for his painstaking criticism of the English.

JOHN B. EKELEY.

INTRODUCTION

TO THE TEACHER

To use this book, especially Parts I. and II., to the best advantage, the order of the experiments, in most cases, should not be changed. Great care has been taken to arrange them so that they follow one another in logical order. If there is a lack of time to do them all, some (the teacher will easily see what ones) can be omitted without destroying the continuity of the whole.

Students should be examined by means of recitations or individual examinations after each ten experiments of Part I. At St. Paul's, each student meets the instructor for a "quiz" as soon as ten consecutive experiments have been finished. With Part II., however, it seems best to supplement the experimental work with frequent recitations.

TO THE STUDENT

Procure a blank book of about one hundred and fifty pages, with durable covers (leather back and corners). Begin your notes on the sixth page, leaving the first five pages blank for a future "table of contents." On the left-hand page, record with a hard pencil the notes on your observations as taken *at the time of the experiment*. Do not fall into the habit of writing your original notes on slips of paper, but record them immediately in your note book. When your

original notes have been criticised by your instructor, study them carefully. After you are satisfied that you thoroughly understand the experiment, write out in ink on the right-hand page in your best English a detailed description of the experiment and the conclusions that you have reached.

Before attempting an experiment, study carefully the description as given in the text, so that you are sure you understand what you are going to do, and what the object of the experiment is.

Keep your laboratory desk and apparatus clean. You will often be tempted not to do this, but you will get the best results from your work, and get the most enjoyment from it, if you avoid slovenly habits of experimentation.

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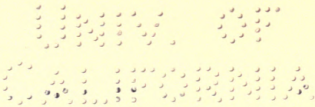
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PART I.

PREPARATION AND PROPERTIES OF ELEMENTS AND COMPOUNDS



AN ELEMENTARY EXPERIMENTAL CHEMISTRY

PART I.

PREPARATION AND PROPERTIES OF ELEMENTS AND COMPOUNDS

EXPERIMENT 1

Three Conditions of Matter

TAKE a piece of ice. Note its properties. After crushing, heat in a 100 cc. flask. When all is melted, note its properties. Heat the water obtained to boiling. Does anything escape from the mouth of the flask? If so, state its properties. What are the three conditions of matter?

EXPERIMENT 2

Physical and Chemical Changes

a. Dissolve a little sugar in water. Evaporate until all the water is gone. What remains?

b. Heat a little sugar in a porcelain dish. Examine the residue. Is it still sugar?

4. AN ELEMENTARY EXPERIMENTAL CHEMISTRY

c. Heat an iron wire in a flame. Examine it again after it is cool. Is it still iron?

d. Have ready a few small iron nails. Pour 10 cc. of water into a 100 cc. flask, and then add 5 cc. of sulfuric acid. In diluting sulfuric acid, always pour the acid into the water. If you pour the water into the acid, it will spatter. *Dangerous.* Place the nails in the diluted acid; and, after all



action has ceased, filter, evaporate, and allow to crystallize. Is the resulting product iron?

State what you conceive to be the difference between physical and chemical change.

EXPERIMENT 3

Mechanical Mixture and Chemical Compound

An element is a substance that has not been separated into two or more dissimilar substances.

Sulfur and iron are elements.

a. Mix equal portions of flowers of sulfur and fine iron filings. Can you distinguish the particles of each element on looking at the mixture through a strong magnifying glass? Pass a magnet through the mixture, and tap it lightly on some solid body. Place a little of the mixture in a test tube, and add a little carbon disulfid. After shaking well for a few moments, pour off the liquid and examine the residue. Is it iron? Evaporate the liquid. What remains?

b. Place the remainder of the mixture in a small test tube, and hold the bottom of the tube in the flame of a Bunsen burner. After the action has ceased, break the

tube, and examine the contents by the same tests as before heating. Is it iron? Is it sulfur?

State what you conceive to be the difference between a mechanical mixture and a chemical compound.

EXPERIMENT 4

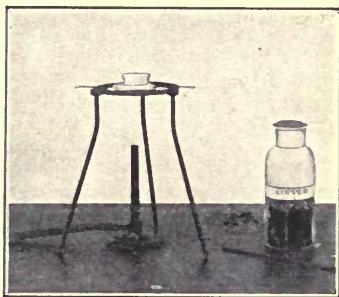
Copper and Copper Oxid

a. Examine some copper in the form of wire and sheet. State as many of its properties as you can; i.e., color, hardness, luster, weight, tenacity, fusibility, volatility, etc.

b. Place about 5 grms. of fine copper filings in a porcelain crucible, and weigh carefully to 1 c. gm. Heat over a Bunsen burner, stirring occasionally with a clean iron rod. After it is cool, weigh again.

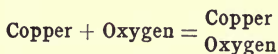
Let us explain the increase in weight. Since matter is the only thing that has weight, some kind of matter must have been added to the contents of the crucible. It has not been added

to the crucible itself; for, by trying, you would find that the weight remains unchanged. This matter could not have come from the heat, since heat is not material; so it must have come from the air. This something chemists call *oxygen*. Examine the contents of the crucible; you will see that it is no longer copper. A chemical change has taken place, and a new substance called *copper oxid* has



been formed. It is thus named to indicate the chemical union of copper and oxygen.

When elements unite chemically, the resulting substance is called a compound. The process of combining two or more elements chemically is called *synthesis*. We may express what has taken place by the following:

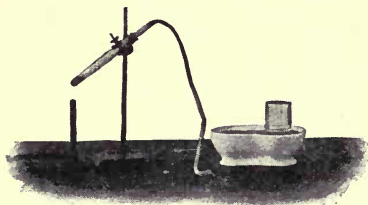


Understand that, when the names of elements are written one under another, the combination signifies that the elements are chemically combined.

EXPERIMENT 5

Mercury

a. Properties. Examine a small quantity of the element mercury. State as many of its properties as you can distinguish. In heating mercury, be careful not to breathe any of the vapor.



By heating mercury a long time in a flask, it is possible to obtain a red powder that

weighs more than the original mercury. This red powder is called *mercury oxid*, and is formed by the union of mercury and oxygen just as copper oxid is formed by the union of copper and oxygen.

b. Analysis of Mercury Oxid. Place a small quantity of mercury oxid in a hard glass tube of about 15 mm. diameter

and closed at one end. Fit the tube with a one-holed stopper through which extends a glass exit tube of about 4 mm. bore. By means of a piece of rubber tubing, connect this with a glass delivery tube. Clamp the hard glass tube on a standard so that it is in a slightly inclined position. Have ready a pneumatic trough with a small beaker full of water inverted on the shelf. Now heat the mercury oxid gently. Gradually increase the heat, and allow the bubbles to escape under the inverted beaker of water. What forms on the sides of the tube? If mercury oxid is composed of mercury and oxygen, what is the substance in the inverted beaker?

This process, the separating of a compound into two or more elements, is called *analysis*.



EXPERIMENT 6

Oxygen

a. Properties. Examine the substance in the jar made in the last experiment. Place a ground-glass plate under the jar, and remove it from the water. Has the gas formed color, odor, or taste? Light a pine splinter, blow out the flame so that a spark remains, and plunge the splinter into the jar. What happens? What is the substance present in the air, that makes fuel burn? Why is it that a fire is smothered when air is kept away from it?

b. Preparation from Manganese Dioxid. Perform an experiment similar to Experiment 5 *b*, using manganese dioxid (composed of manganese and oxygen) instead of mercury oxid. Is a metal left as in experiment 5 *b*?

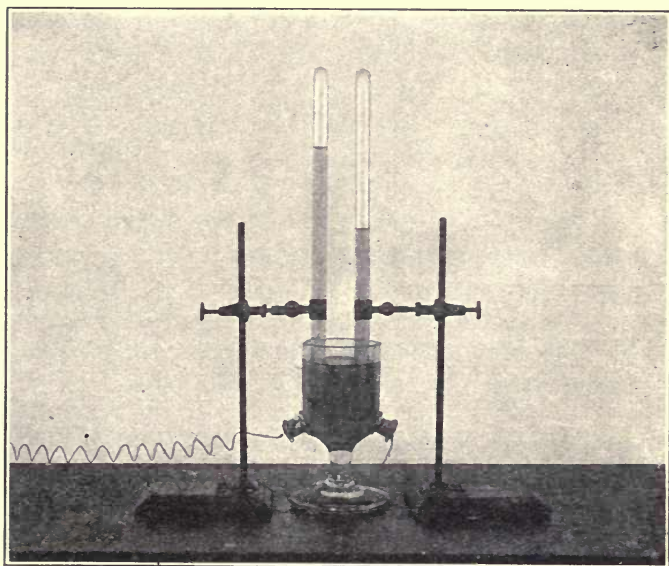
c. Preparation from Potassium Chlorate. Perform a similar experiment, using potassium chlorate instead of mercury oxid. Heat gently at first, beginning at the top of the chlorate, and be careful to keep the hands from under the tube in case it should break; for melted potassium chlorate makes a serious wound. After the action has ceased, dissolve the remaining substance in hot water, and evaporate to dryness. Preserve this residue, and, when you have performed Experiment 25 *c*, compare.

d. Preparation from a Mixture of Potassium Chlorate and Manganese Dioxid. Perform a similar experiment, using 8 grms. of potassium chlorate and 2 grms. of manganese dioxid. Mix the two thoroughly. Begin heating at the top of the mixture. When action has ceased, add hot water and transfer to a beaker. Be sure to get all the residue. Weigh a filter paper carefully. Filter the hot liquid, being careful to throw it all and the black residue upon the filter paper. Wash with hot water from a wash bottle five or six times. Dry the filter in an oven at a temperature of about 100° . Weigh again. Subtract the weight of the filter paper from the weight of the filter paper and black residue. How does the result compare with the original weight of the manganese dioxid? Has the manganese dioxid undergone any change chemically? Evaporate the filtrate (the liquid that ran through the paper) to dryness in a weighed porcelain dish. Weigh, and account for the loss of weight of the chlorate.

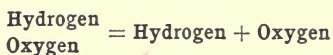
e. Preparation by Electrolysis of Water. Have ready a vessel fitted with two platinum electrodes. Into this pour sufficient water acidified with sulfuric acid (1 part acid to 20 of water),* so that it will somewhat more than cover the electrodes. Fill two glass tubes graduated in cubic centi-

* The acid is only added to aid the water in conducting the current.

meters with some of the water, and invert them in the vessel so that the mouth of each tube covers an electrode. Place the vessel in circuit with a battery of two bichromate cells. Note the evolution of gas at each electrode. After a few minutes, disconnect the cells, and measure the relative volumes of the gases. Test the lesser one with a glowing



splinter. What is it? Apply a lighted match to the greater, and note the result. Here we have a new gas, which chemists have been unable to decompose into anything else. It has been named *hydrogen*. We now have experimental proof as to the composition of water.



Remark. If a silent discharge of electricity is passed through dry oxygen, the gas will decrease in volume one-third, and will acquire greater chemical activity. Some change, then, has occurred in the gas without the addition of any other kind of matter. When elements can thus be changed by any means, they are said to exist in *allotropic forms*. This second form of oxygen is called *ozone*. It has a faint peculiar odor, often noticed near electric machines when operating, or just after a discharge of lightning.

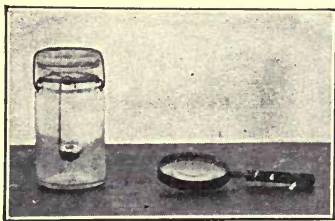
EXPERIMENT 7

Phosphorus

Phosphorus occurs in two allotropic forms, the yellow and the red. The yellow must be stored under water, as it ignites at a temperature of 40° in the air.

a. Properties. Examine both red and yellow phosphorus. Be very careful in handling the yellow variety, since it ignites so easily, and a burn from it heals with difficulty. *Never*

handle it with the fingers, and always cut it under water. State as many properties as you can of each kind.

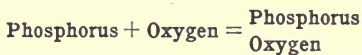


b. Preparation of Phosphoric Oxid by Burning Phosphorus in air. Have ready a quick-sealing fruit

jar. The rubber washer should be greased with vaseline. On a support in the jar, hang a deflagrating spoon containing a piece of ignited* phosphorus about the size of

* The phosphorus may be placed in the jar, and then ignited by means of a strong lens.

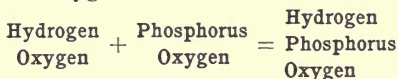
a pea. Close the jar air tight. Note what forms in the jar. This product is called *phosphoric oxid*. Open the jar under water. Was any of the air in the jar used up? Does the oxid dissolve in water?



Remark. Another oxid of phosphorus exists, called *phosphorous oxid*. It is formed when phosphorus is burned in an insufficient supply of air. It is a white powder with an odor resembling garlic. When heated in the air, it becomes the oxid you have just made; this shows that it differs from phosphoric oxid, in that it contains less oxygen.

c. Preparation of Phosphoric Oxid by burning Phosphorus in Oxygen. Perform a similar experiment, using a jar of oxygen instead of air. In this case, it is better to ignite the phosphorus in the jar by means of a burning glass. How does the product compare with that formed in *b*?

d. Phosphoric Acid. Dissolve the oxid prepared in *c* in a small amount of water. Taste a drop of the solution. Also note its effect upon blue litmus paper. This product is called *phosphoric acid*; and, since it is made of phosphoric oxid and water, it can contain only the elements hydrogen, phosphorus, and oxygen.



Remark. When phosphorus is burned in an insufficient supply of moist air, another acid of phosphorus is formed which has a faint odor like garlic. This is called *phosphorous acid*. It differs from phosphoric acid in that it contains less oxygen. Note the signification of the endings *ic* and *ous*.

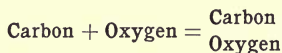
Definition of an Acid. For the present it will be sufficient to define an acid as a compound of hydrogen, or else of hydrogen and oxygen, with a non-metallic element, which compound has usually a sour taste and turns blue litmus paper red.

EXPERIMENT 8

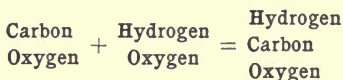
Carbon

a. Properties. Examine, and state the properties of, the following allotropic forms of carbon: charcoal, gas carbon, bone black, graphite, and soot.

b. Carbon Dioxid. Burn a piece of charcoal in a deflagrating spoon in a jar of oxygen, as in Experiment 7 *b*. The charcoal must be well ignited before being placed in the jar. Open the jar under water. Does the cover stick? What can you say of the volume of the resulting carbonic oxid compared with the original volume of oxygen? Close the jar, allowing a little water to enter, and shake vigorously. Open again under water. Is carbonic oxid soluble in water? In the same way, prepare another jar of the gas. Plunge a lighted splinter into the gas. Does the gas burn? Does the splinter continue to burn? The name *carbon dioxid* is given to this gas for reasons that will appear later.



c. Carbonic Acid. Taste the solution formed in *b*, and test it with blue litmus paper. Here we have a second acid, which has been named *carbonic acid*. Transfer to a test tube and heat to boiling. Taste, and test with litmus paper again. Is carbonic acid a stable compound, i. e., is it hard to decompose? What has become of the carbon dioxid?



Remark. In Experiment 7 *b*, and 8 *b*, we have observed the phenomenon of combustion. *Combustion* is the chemical union of substances, accompanied by the evolution of light and heat. The substances that unite may be any substances whatever, but ordinarily we apply the term to the union of substances with the oxygen of the air. The substance that unites with the oxygen is called the *combustible*, while the oxygen is said to *support combustion*. The ease with which elements unite with oxygen varies. Some, such as phosphorus, require only a slight rise in temperature, while others, like carbon, require a comparatively high one. The temperature at which a substance takes fire in air is called its *kindling temperature*. This temperature is constant for any particular substance.

EXPERIMENT 9

Sulfur

a. Properties. After examining, state as many properties as you can of roll sulfur. Place about 10 grms. in a test tube, and heat over a Bunsen burner, heating gently at first and gradually increasing the temperature. Note carefully the changes that the sulfur passes through, especially in regard to its color and consistency. Note also the ease with which sulfur takes fire.

b. Allotropic Forms of Sulfur.

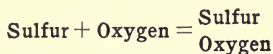
1. In a beaker put about 60 grms. of roll sulfur, and heat gently until all is melted. When all is melted, allow it to

cool. As soon as crystals have just covered the surface of the liquid, break a hole in the crust, and quickly pour out the melted sulfur. Note the shape and color of the crystals that remain in the beaker. Allow it to stand for a few days, and examine again.

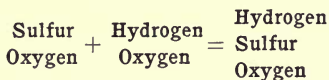
2. Pulverize a small lump of roll sulfur, and shake it up with carbon bisulfid in a test tube. Keep away from a flame, since carbon bisulfid is very inflammable. When it is dissolved, pour it into an evaporating dish and allow it to crystallize. Note the color and shape of the crystals formed.

3. In a test tube, heat to boiling about 15 grms. of roll sulfur, and quickly pour the liquid into a beaker of cold water. State the properties of the product. Allow it to stand for a few days, and examine again.

c. Sulfur Dioxid. Burn a small amount of sulfur in a deflagrating spoon in a jar of oxygen. Note the color of the flame. After action ceases, open under water. Does the cover stick? What can you say of the volume of the gas compared with the original volume of oxygen? Get its odor and color, if any. Allow a little water to enter the jar, seal again, and shake. Open again under water. Does the cover stick? Is the gas soluble in water? We shall call this gas sulfur dioxid for reasons that will appear later.

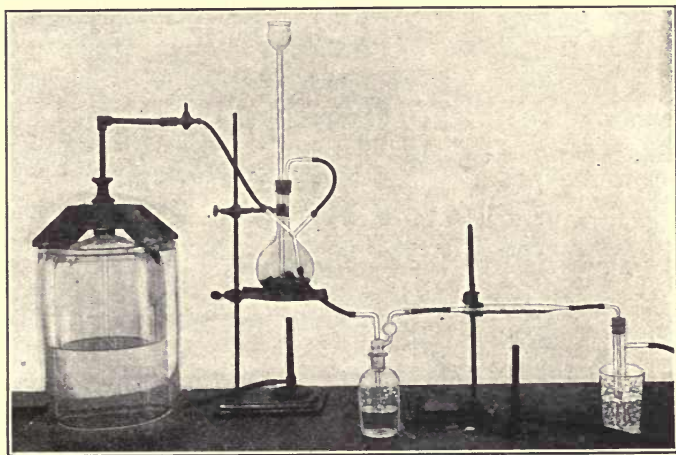


d. Sulfurous Acid. Taste the solution made in *b*, and test it with blue litmus paper. This is called *sulfurous acid*.



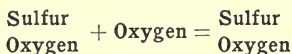
e. Sulfur Trioxid. Have ready a gas holder full of oxygen. Apply to the instructor for an apparatus for obtaining

a steady stream of sulphur dioxid. (See *g.*) By means of rubber tubing and a glass Y tube, allow the oxygen and sulfur dioxid to come together. In order to dry the mixed gases, pass them by means of rubber tubing through a catch bottle containing concentrated sulfuric acid, which has the property of extracting moisture from gases as they bubble through it. Then pass the dry gases through a hard glass tube, about 15 cm. long and 1 cm. bore, containing platinized

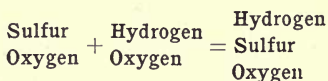


asbestos. Do not pack the asbestos in the tube, but place it in loosely. Have the tube containing the asbestos attached to a test tube fitted with a two-holed stopper containing an entrance and an exit tube. Let the entrance tube extend down to within two centimeters of the bottom. Pack the test tube in a mixture of ice and salt. Now heat the asbestos to redness. The red-hot platinized asbestos simply aids in the chemical union of the two gases. After three or four minutes, remove the test tube from the ice. What is

the nature of the crystals formed? Note especially the effect of warmth upon them. This product is called sulfur trioxid. Since it is formed by the chemical union of sulfur dioxid and oxygen, it differs from sulfur dioxid in that it contains more oxygen.



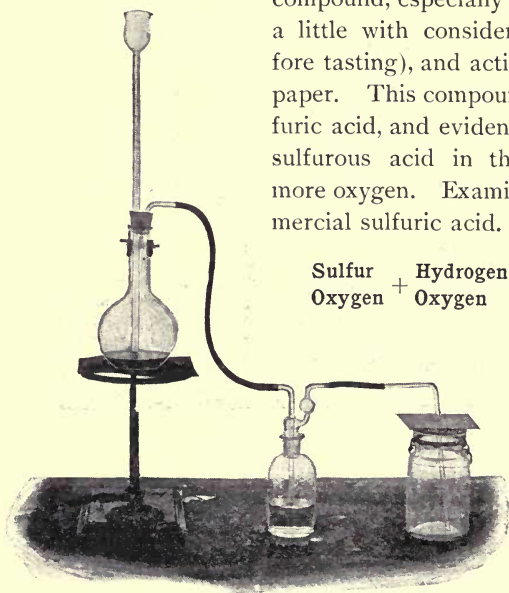
f. Sulfuric Acid. Allow a few drops of water to fall upon the oxid made in *e*. Get the properties of the resulting compound, especially its taste (dilute a little with considerable water before tasting), and action upon litmus paper. This compound is called sulfuric acid, and evidently differs from sulfurous acid in that it contains more oxygen. Examine a little commercial sulfuric acid.



g. Sulfur Dioxid from Sulfuric Acid and Copper.

In a 250 cc. flask, place about 50 grms. of sheet copper clippings.

Fit the flask with a two-holed stopper containing an exit tube and a funnel tube that extends down to the copper. Pour concentrated sulfuric acid into the flask till it somewhat



more than covers the end of the funnel tube. Heat with a Bunsen burner, and collect a jar full of the gas by means of a glass tube passing to the bottom of the jar. Note that it is the same gas as you obtained on burning sulfur in oxygen, i.e., sulfur dioxid.

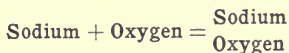
EXPERIMENT 10

Sodium

The element sodium must be stored under naphtha, to preserve it. When you wish to use any, apply to the instructor for it. Never handle it with the fingers. Use forceps.

a. Properties. Examine a piece of clean sodium. Get its properties, — hardness, color, etc. Is it a metal? Make a fresh cut on it with a knife, and note the appearance of the clean surface for a few moments. Throw a small piece upon water, and note the result.

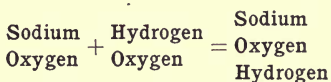
b. Sodium Oxid. In a crucible, heat a piece of clean sodium about the size of a pea. As it burns, note the color of the flame. Get the properties of the resulting sodium oxid.



c. Sodium Hydroxid. When the oxid made in *b* is cooled, allow a few drops of water to fall on it. Note the sputtering sound. Put a drop of the solution on your finger, and note the greasy feeling. Test the liquid with red litmus paper. Evaporate it to dryness, and allow it to stand for some time, after which note that moisture has collected upon the substance.

Since sodium oxid has united with water, the resulting compound must be composed of sodium, oxygen, and hydro-

gen. It is called sodium hydroxid or sodium hydrate. This is the first of a class of compounds called *bases*, which you will study.



Definition of a Base. For the present it is sufficient to define a base as a compound of a metallic element with oxygen, or with oxygen and hydrogen. Bases that turn red litmus paper blue are called *alkalis*.

d. Sodium Amalgam. Weigh out 20 grms. of mercury, and place it in an iron pan under a hood. Heat it to about 200°. The temperature may easily be observed by means of a thermometer. Take 2 grms. of clean sodium, and by means of a pair of long forceps drop it into the mercury. Step back instantly, since mercury and sodium unite with great violence, and poisonous vapors of mercury are evolved. The product is called sodium amalgam.

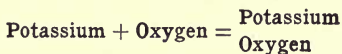
EXPERIMENT 11

Potassium

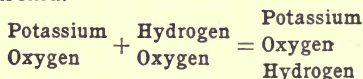
Potassium is stored in a way similar to sodium, and must be handled with equal care.

a. Properties. Note the properties of potassium as you did those of sodium. Compare the metal with sodium.

b. Potassium Oxid. Perform an experiment with a piece of potassium, corresponding to Experiment 10 *b*. Note the color of the flame.



c. Potassium Hydroxid. Perform an experiment with potassium corresponding to Experiment 10 *c*. How does the compound here made differ in chemical composition from that made in Experiment 10 *c*? Here we have another compound belonging to the class of bases. It is called potassium hydroxid.

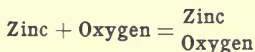


EXPERIMENT 12

Zinc

a. Properties. Examine, and note the properties* of, the metal zinc in its various forms, — sheet, stick, granular, and dust.

b. Zinc Oxid. Heat a small piece of zinc in a crucible, stirring with an iron rod. Note the color of the flame as the zinc burns, the color of the oxid when hot and when cold, and the peculiar woolly appearance that it assumes.



c. Zinc Oxid and Water. Try the effect of water on a little oxid of zinc, and test it with litmus paper turned red with carbonic acid. Filter, and evaporate the filtrate. Is zinc oxid soluble in water? Does zinc oxid form an acid or a base?

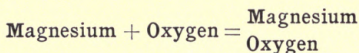
* Zinc is peculiar in that its physical properties vary considerably with heat. At the ordinary temperature, zinc is rather brittle, but, if heated to 100°, it becomes malleable and can be rolled into sheets. At 205° it becomes so brittle that it can be powdered. It melts at 400°, and boils at about 1000°. Granular zinc is formed when the molten metal is allowed to fall drop by drop into cold water. Zinc dust is obtained when the vapors of boiling zinc are suddenly condensed in the absence of air.

EXPERIMENT 13

Magnesium

a. Properties. Examine the element magnesium both in the form of ribbon and powder (flash light powder). State its properties.

b. Magnesium Oxid. Oxidize a little magnesium in a crucible, noting the color of the oxid when hot and when cold. How could you distinguish magnesium oxid from zinc oxid?



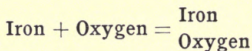
c. Magnesium Oxid and Water. Treat magnesium oxid with water as you did zinc oxid in Experiment 12 *c*. Does magnesium oxid form an acid, or a base?

EXPERIMENT 14

Iron

a. Properties. Examine, and state the properties of, iron both in the form of nails and in the form of powder (iron by hydrogen).

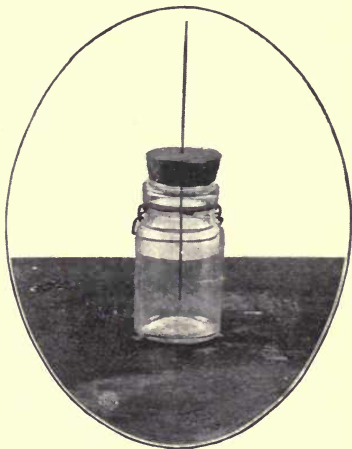
b. Iron Oxid. (Burning iron in air.) Place about 10 grms. of "iron by hydrogen" in a crucible, and heat, stirring occasionally. The resulting compound is evidently iron oxid.



c. Iron Oxid and Water. Try the effect of water upon iron oxid.

d. Iron Oxid. (Burning iron in oxygen.) Fill a jar with

oxygen gas. Pierce a hole through a flat cork large enough to cover the mouth of the jar. Through this hole, insert a watch spring from which the temper has been taken by heating in the flame of a Bunsen burner. Around one end of the spring, wind a little cotton string, and soak this in melted sulfur. Ignite the sulfur, and plunge it into the jar of oxygen, making the cork cover the mouth of the jar. As the spring is consumed, feed it through the hole until action ceases. What compound is formed?



EXPERIMENT 15

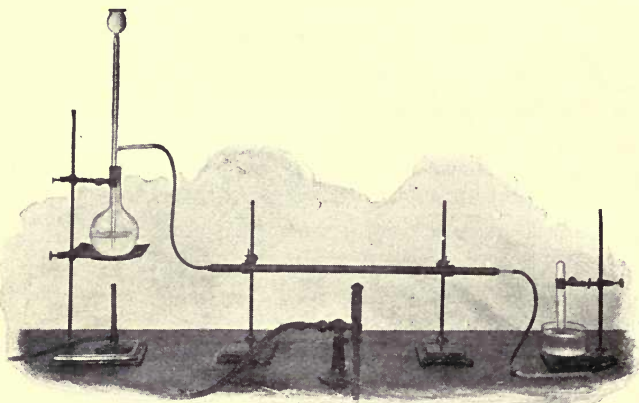
Hydrogen

a. Hydrogen by Electrolysis of Water. Properties. Make hydrogen again by electrolysis of water, as in Experiment 6 *e*. Note its properties.

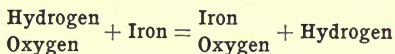
b. Hydrogen from Steam and Hot Iron. Place about 25 or 30 grms. of *clean* iron filings in the middle of a piece of half-inch gas pipe about two feet long. Be careful that the filings do not stop up the pipe. Fit the ends with one-holed stoppers containing glass tubes about 5 mm. bore. To one end of the pipe connect a 500 cc. flask about half full of water, and to the other a rubber tube fitted with a glass leading tube. Heat the gas pipe red hot with a blast lamp, and then boil

the water in the flask. Collect the gas that issues from the leading tube in test tubes over water. Compare the gas with the hydrogen obtained in *a*. Be careful to have no flame near the leading tube, since hydrogen mixed with air is an explosive compound.

Let us see what has taken place. We know from Experiment 6 *e* that water is composed of oxygen and hydrogen,



and from Experiment 14 *b* we know that iron will unite chemically with oxygen. Examine the substance left in the gas pipe, and recognize it as iron oxid.

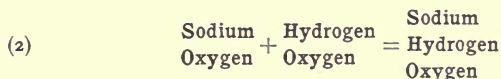
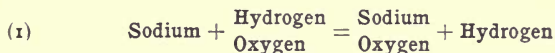
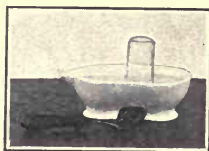


c. Hydrogen from Sodium and Water. In a cage made of fine wire gauze and fitted with a handle, place a small piece of metallic sodium about the size of a pea. Have ready a pneumatic trough with an inverted beaker full of water on the shelf. Hold the cage containing the sodium under water. After a few bubbles have escaped, hold it under

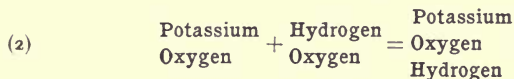
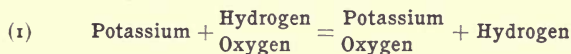
the beaker and collect the gas evolved. When action has ceased, recognize the gas as hydrogen.

Into a small beaker full of water, throw a few small pieces of sodium, and test the liquid afterwards with red litmus paper.

Sodium has so strong an attraction for oxygen that it takes oxygen out of water, leaving the hydrogen. The resulting oxid unites with water and forms the hydroxid. If there is an insufficient supply of water, the sodium bursts into the characteristic yellow flame. To show this, place a small piece of sodium on a wet piece of filter paper, and note the result.



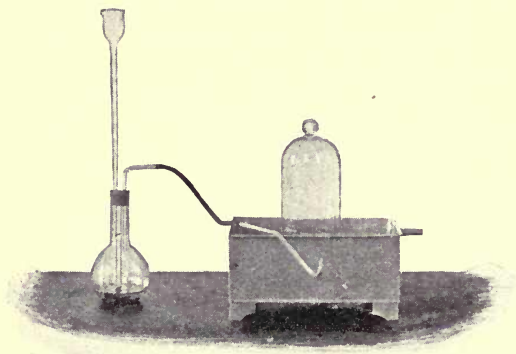
d. Hydrogen from Potassium and Water. Perform an experiment similar to *c*, using potassium instead of sodium.



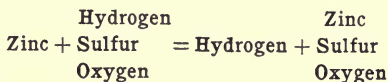
e. Hydrogen from Zinc and Sulfuric Acid. In a 500 cc. flask fitted with a two-holed stopper containing an exit and a funnel tube, place about 100 grms. of granular zinc. Add enough sulfuric acid (1 part acid and 5 parts water) to cover the zinc. Catch the gas evolved over water, and recognize as hydrogen. Be sure to have no flame near the flask.

After no more zinc will dissolve, filter the contents of the flask; then, after the solution has evaporated somewhat, allow it to crystallize. Note the properties of the crystals. Place a dry crystal in a glass tube closed at one end, and heat gently in the flame of a Bunsen burner. Notice the water that collects on the upper, cool part of the tube. Does the crystal remain intact? This water, then, seems in some way to be necessary to the crystal, and is called "water of crystallization."

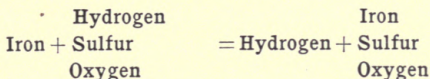
Since sulfuric acid is composed of hydrogen, sulfur, and oxygen, and since in the experiment the zinc disappears and



hydrogen is evolved, it is evident that the zinc takes the place of the hydrogen in the sulfuric acid. The crystals are therefore composed of zinc, sulfur, and oxygen, together with the water, which holds the combined elements in crystalline form. When this water is driven off, of course only zinc, sulfur, and oxygen are left, in the form of a white powder. This is called zinc sulfate.



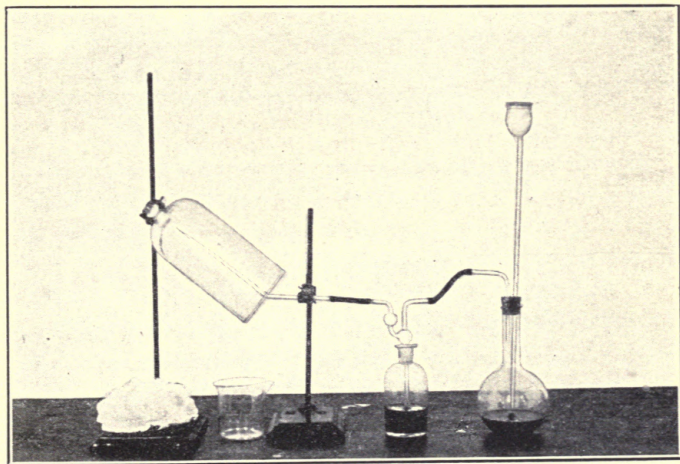
f. Hydrogen from Iron and Sulfuric Acid. Perform a similar experiment, using iron nails instead of zinc. What is the color of the crystals obtained? Of what are they composed? What would you naturally name them?



EXPERIMENT 16

Hydrogen and Oxygen

a. Synthesis of Water by Burning Hydrogen in Air. Set up an apparatus as in Experiment 15 *e*. To dry the hydrogen,



pass it through a catch bottle containing concentrated sulfuric acid, and have the delivery tube drawn to an opening of about 1 mm. diameter. After the gas has been escaping

for a few minutes, collect a test-tubeful over water. If, on ignition, no explosion is heard, it is safe to light the hydrogen as it escapes from the apparatus. Never light hydrogen from a generator without taking this precaution. Hold a dry bell jar over the flame, and keep the jar cool with a cloth wrung out in cold water. Notice the water forming on the inside of the jar. Here we have further proof of the composition of water. In Experiment 6 *c*, we proved by analysis that water is composed of hydrogen and oxygen. Here the same thing is proved by synthesis.

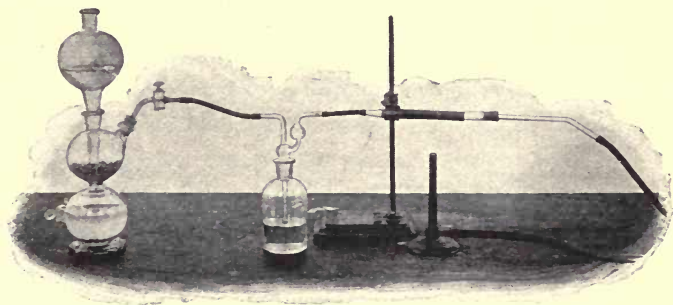
Remark 1. Water is also formed by the oxidation in the air of substances containing hydrogen. Ignite a piece of wood, and hold the flame near a cold glass plate. Note the moisture condensed upon the plate.

By the slow oxidation in the human body of substances containing hydrogen, water is formed. Breathe upon a cold glass plate, and note the moisture condensed upon it. For a second product of oxidation in the human body, see Experiment 20 *d*.

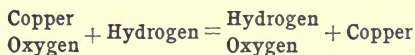
Remark 2. If the earth were covered with an atmosphere of hydrogen instead of oxygen, a stream of oxygen issuing from a jet would burn, if ignited, just as the stream of hydrogen burns in this experiment.

b. Reduction of Copper Oxid by Hydrogen. In a combustion tube, place about 5 grms. of copper oxid. Have ready a hydrogen generator connected with a catch bottle containing concentrated sulfuric acid, and attach this to one end of the combustion tube. To the other end, attach a dry test tube fitted with a two-holed stopper containing entrance and exit tubes, and place this in a beaker of cold water. Allow the hydrogen to pass through the apparatus for a few moments, to drive out the air; then, when it is safe, heat the copper oxid. Notice that, after action begins, the oxid glows brightly.

When all is finished, allow the contents of the tube to cool in the stream of hydrogen. Meanwhile disconnect the test tube, and note the water collected in it. When the combustion tube is cool, remove the contents and recognize as me-



tallic copper. Such a process as this, i.e., the taking away of oxygen from a substance, is called *reduction*. Adding oxygen to a substance is called *oxidation*. Which substance in this case is reduced? Which oxidized?



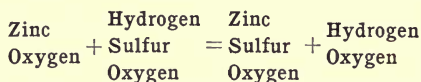
EXPERIMENT 17

Zinc and Magnesium Oxids and Sulfuric Acid

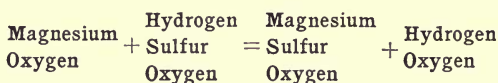
Review Experiment 15 *b, c, d, e, and f.*

a. Zinc Oxid and Sulfuric Acid. Add zinc oxid to dilute sulfuric acid (1 part acid to 5 of water) in a flask, until there is no further action. Notice that no hydrogen is given off. Evaporate the solution, crystallize, and recognize the crystals as the same as those obtained in Experiment 15 *e*, i.e., zinc sulfate. Since no hydrogen is given off, we are com-

pelled to conclude that it united with the oxygen of the zinc oxid, and formed water.



b. Magnesium Oxid and Sulfuric Acid. Perform a similar experiment, using magnesium oxid instead of zinc oxid. Name the crystals.



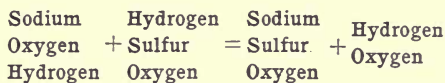
EXPERIMENT 18

Neutralization

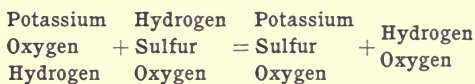
a. Sodium Hydroxid and Sulfuric Acid. Take about 5 cc. of concentrated sulfuric acid diluted with about 10 cc. of water. Dissolve about 10 grms. of sodium hydroxid in water. Add the hydroxid to the acid, until the solution shows neither an alkaline nor acid reaction upon litmus paper. Evaporate and crystallize it. Do the crystals contain water of crystallization? (See Experiment 15 *e*.) Allow a dry crystal to lie in the air for some time, and examine it. What has happened to it? This phenomenon is called *efflorescence*, and the crystal is said to *effloresce*.

Since sodium hydroxid is composed of sodium, oxygen, and hydrogen, and sulfuric acid is composed of hydrogen, sulfur, and oxygen, and since we have learned (see Experiment 15 *e, f*) that the tendency of metals is to force hydrogen out of compounds, it is evident that the sodium of the sodium hydroxid has united with the sulfur and oxygen

of the acid, leaving the hydrogen of the acid to unite with the oxygen of the hydroxid. The resulting crystals we call sodium sulfate.



b. Potassium Hydroxid and Sulfuric Acid. Neutralize potassium hydroxid with sulfuric acid as in *a*. Evaporate and crystallize. We name these crystals potassium sulfate.



Remark. A compound obtained by replacing the hydrogen of an acid by a metal, we call a *salt*. We may now remodel our definitions of acids and bases. Our full definitions will then be:—

An *acid* is a compound of one or more non-metallic elements with hydrogen, or with hydrogen and oxygen, all or part of whose hydrogen may be replaced by a metal.

A *base* is a compound of a metallic element with oxygen, or with oxygen and hydrogen, the metal of which is capable of replacing the hydrogen of an acid.

We see that, in neutralizing an acid with a base, we shall always obtain a salt and water.

Acid, Normal, and Basic Salts. An *acid salt* is a salt formed by replacing only part of the hydrogen of an acid by the metal of a base.

A *normal salt* is a salt formed by replacing all the replaceable hydrogen of an acid by a metal of a base.

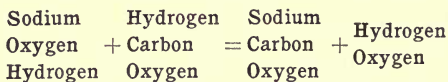
A *basic salt* is a salt formed by replacing all the replaceable hydrogen of an acid by a metal of a base, and, in addition, causing more of the base to unite with the compound.

EXPERIMENT 19

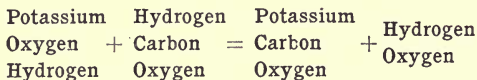
Carbonic Acid and Carbon Dioxid with Sodium and Potassium Hydroxid

a. Dissolve about 1 gram. of sodium hydroxid in 20 cc. of water, and add from this to a solution of carbonic acid, until it no longer turns blue litmus paper red. Evaporate and crystallize. Do the crystals contain water of crystallization? Do they effloresce?

Arguing as in Experiment 18 *a*, we come to the conclusion that the salt formed is made up of the elements sodium, carbon, and oxygen, chemically combined. The crystals are called sodium carbonate or sal soda.



b. Perform a similar experiment, using potassium hydroxid instead of sodium hydroxid.



c. Pass carbon dioxid through solutions of sodium and potassium hydroxids, and obtain the same salts.

EXPERIMENT 20

Calcium

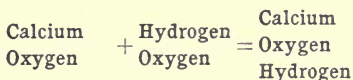
On account of its expense and the difficulty of preserving the element calcium, it is improbable that there will be an opportunity to notice its properties. Suffice it to say that it

is a yellowish metal, which may be kept for some time in dry air, but which in the presence of water, oxidizes with the evolution of hydrogen, forming calcium hydroxid. While the element itself is unimportant, its compounds must be carefully studied. The oxid we know familiarly as quick lime.

a. Properties of Calcium Oxid. Examine a lump of calcium oxid, and note its properties.

b. Calcium Hydroxid. Place a lump of calcium oxid about the size of a walnut in a porcelain dish. Allow water to drip upon this as long as any is absorbed. Notice the evolution of heat. Test the product with moist litmus paper.

The water must have united with the calcium oxid, thus forming calcium hydroxid; just so in Experiment 10 c, water united with sodium oxid and formed sodium hydroxid.

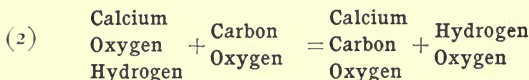
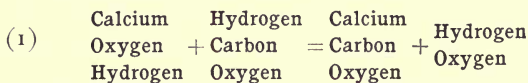


c. Lime Water. Put a little of the hydroxid made in *b* in a beaker of cold water. Stir, and, after allowing it to settle, pour off the liquid. Fill the beaker again with water, and, after stirring vigorously, filter it into a bottle that can be tightly stoppered. Place a little of the liquid on a watch glass, and evaporate it. Is calcium hydroxid soluble in water? The liquid in the bottle is commonly called lime water.

d. Lime Water and Carbonic Acid. In a test tube, place 1 cc. of lime water. Add 1 cc. of carbonic acid, and note the precipitate formed. Now add more carbonic acid, until the solution becomes clear. Boil until the precipitate returns.

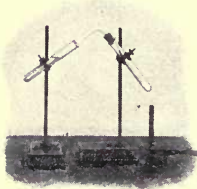
From a generator (described in Experiment 26) pass carbon dioxid into 1 cc. of lime water, until the precipitate which at first appears is dissolved. Boil as before.

We obtain a salt just as in the experiment with sodium hydroxid and carbonic acid ; but, in this case, the salt is calcium carbonate. This salt is evidently soluble in excess of carbonic acid, and reappears in the solution when the carbonic acid is decomposed by heat (see Experiment 8 c). This action, then, can be used as a test for the presence of carbon dioxid, or, conversely, as a test for the presence of calcium hydroxid. See whether or not your breath* contains carbon dioxid.



EXPERIMENT 21

Analysis of Marble (Calcium Carbonate)

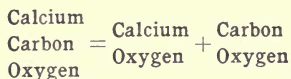
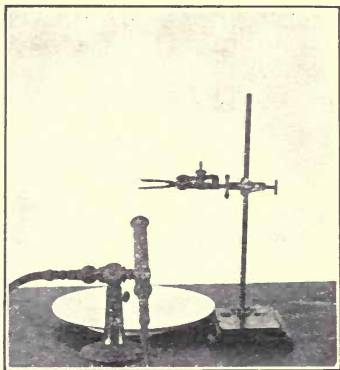


Take a piece of marble about the size of a hazel nut, and grind it to a fine powder in a mortar. Place this in a hard glass tube closed at one end. Fit the tube with a one-holed stopper containing an exit tube extending a centimeter beneath the surface of a little lime water in a test tube. Heat the

* The production of carbon dioxid by the processes of respiration and combustion depletes the supply of oxygen in a room. The oxygen must be renewed from the air outside, and the injurious respiration and combustion products must be removed by proper ventilation. This is especially true of churches, theaters, schoolrooms, and all in-door places where many persons are assembled at the same time.

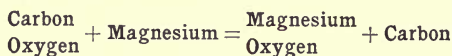
hard glass tube with a blast lamp. After the air has been driven out, notice the effect of the bubbles upon the lime water. Continue heating it until the milkiness of the lime water disappears. Then take away the lime-water tube (before taking away the lamp), and heat it over a Bunsen burner. Does the milkiness return? We therefore know that one constituent of marble is carbon dioxid (see Experiment 20 *d*).

Take another lump of marble about the size of a pea, and, holding with a pair of forceps over a porcelain plate, heat it with a blast lamp for about five minutes. Collect the particles that have fallen on the plate, and, together with what is left of the lump, place them in a very small quantity of water (only enough to moisten them) on a watch glass. Test this with moist litmus and turmeric paper. Now add more water (about 1 cc.), and filter into a test tube. Add a little carbonic acid, and notice the milkiness formed. On adding more carbonic acid, note that it disappears, and reappears on heating. Hence we see that another constituent of marble is calcium oxid (see Experiment 20 *d*), and that marble is composed of calcium, carbon, and oxygen. Its chemical name is calcium carbonate.



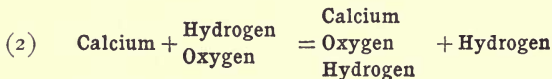
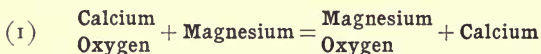
To make a complete analysis, we must take two more steps. First take a jar of carbon dioxid, and, after igniting

a ribbon of magnesium, plunge it into the jar. Note that the magnesium burns to the white oxid, and that carbon is deposited on the sides of the jar.



Secondly take about 5 grms. of calcium oxid, and grind it to a powder in a mortar. Mix with this 2.5 grms. of metallic magnesium powder, place the mixture in an iron crucible, and heat it over a Bunsen burner. While the contents of the crucible are still warm, hold it by means of a pair of long forceps under a beaker full of water and inverted on the shelf of a pneumatic trough. Recognize the gas collected as hydrogen.

Since neither magnesium, magnesium oxid, nor calcium oxid decomposes water with the evolution of hydrogen (see Experiments 13 and 20), it must be that metallic calcium was formed in the crucible and, on being placed in the water, decomposed it, giving hydrogen.



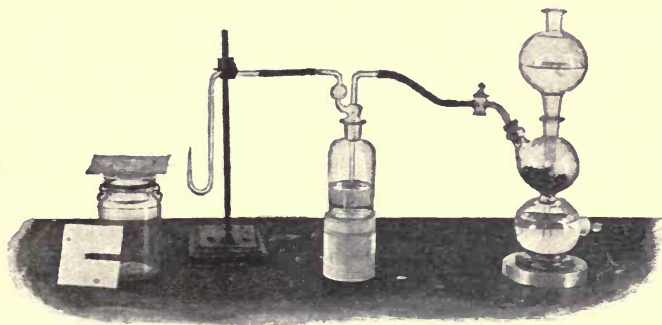
EXPERIMENT 22

Chlorin

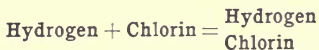
a. Properties. Apply to the instructor for a jar of chlorin. In working with this element, perform your experiments under a hood with a strong draught, since chlorin taken into the lungs is very *dangerous*. Note the color and odor (you can

hardly avoid smelling it). Moisten a little blue and red litmus paper, also a colored flower, if possible, and allow them to hang in the gas for a few moments. Pour about 100 cc. of water into the jar; seal, shake, and open it under water. Is chlorin soluble in water?

b. Hydrogen Chlorid (Hydrochloric Acid) by Burning Hydrogen in Chlorin. Have ready a hydrogen generator (see Experiment 16 a). When the air is driven out, ignite the jet and allow it to burn in a jar of chlorin until no green



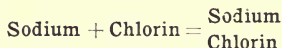
color remains. Blow your breath across the mouth of the jar, and notice that the gas fumes in the presence of moisture. Test with moist litmus paper. Taste the gas by allowing a little to enter the mouth. This is called hydrogen chlorid, or hydrochloric acid gas.



c. Union of Hydrogen and Chlorin by means of Light. To be performed by the instructor. In a dark room with ruby light (a photographic dark-room), fill a 250 cc. flask one-half full of chlorin and one-half full of hydrogen, collected over warm water. Fit a rubber stopper to the flask, and

fasten it securely with wire. Wrap a towel or thick cloth around the flask so that no light can enter. Tie a long string to a corner of the towel, and carry all out of doors. At a safe distance, drag off the towel by means of the string so that the sunlight may strike the flask. What caused the action, and what is formed?

d. Sodium Chlorid. Apply to the instructor for a jar full of *dry* chlorin. Into this, drop a thin slice of metallic sodium about half a square centimeter in area, and seal. Allow it to stand until the next day. Then open the jar, and remove the white compound. Crush it, be *sure* there is no metallic sodium left, and then taste it. What is it? Being formed by the synthesis of sodium and chlorin, it is called sodium chlorid. Note that, in cases where *two* elements unite, the name always ends in *id*.

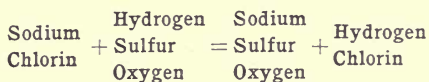


EXPERIMENT 23

Hydrochloric Acid from Sodium Chlorid and Sulfuric Acid

In a 250 cc. flask fitted with a two-holed stopper containing funnel and exit tubes, place about 10 grms. of sodium chlorid. Have the exit tubes connected with a glass tube reaching to the bottom of a fruit jar covered with a piece of cardboard. Into the funnel tube, pour 15 cc. of concentrated sulfuric acid diluted with 4 cc. of water. Pour the acid into the water. Collect two or three jars of the gas. Note that it is the same gas as that obtained by burning hydrogen in chlorin. To one of the jars, add about 25 cc. of water, close and shake it. Open it under water. Is hydrochloric acid soluble in water? Into another jar, pour a little

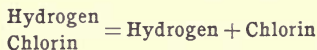
water, shake and test the liquid with blue litmus paper. Dilute the contents of the flask with water after all the gas possible has been allowed to escape. Evaporate and crystallize. Recognize the crystals as the same as those obtained in Experiment 18 *a*, i.e., sodium sulfate.



EXPERIMENT 24

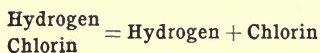
Analysis of Hydrochloric Acid

a. By Electrolysis. Have ready an apparatus like that of Experiment 6 *e*, except that the electrodes are of gas carbon instead of platinum. Fill it with concentrated hydrochloric acid solution. On turning on the current, note the immediate evolution of hydrogen at the negative pole. Since you know that chlorin is soluble in water, you can easily account for the fact that chlorin does not immediately appear at the positive pole. Allow the current to pass through the solution for some time; then, when it has taken up all the chlorin it can hold, hydrogen and chlorin will appear in equal volumes at the negative and positive poles respectively.



b. By Means of Sodium. Fill a *dry* quick-sealing fruit jar with hydrochloric acid gas dried by being passed through concentrated sulfuric acid. Be sure to have the rubber washer well greased with vaseline, so that the jar will be air tight when closed. Prepare about 25 grms. of sodium amalgam (10 parts mercury to 1 part sodium), and powder it in a

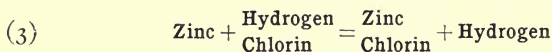
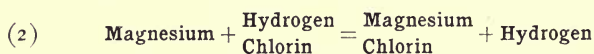
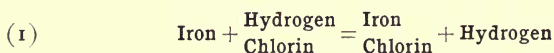
mortar. Drop the powdered amalgam into the jar, seal quickly, shake vigorously, and note the action. Open it under water in a glass or porcelain dish, and, allowing the mercury to fall out, note the volume of the remaining gas. Remove the jar from the water, and recognize the gas as hydrogen. If the experiment is successful, you will have one-half a jar of hydrogen, showing that hydrochloric acid is composed of equal volumes of hydrogen and chlorin.



EXPERIMENT 25

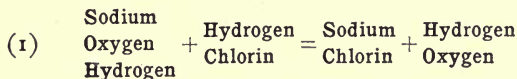
Hydrochloric Acid with Metals, Hydroxids, and Carbonates

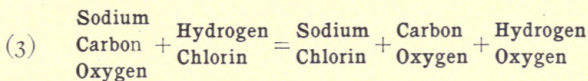
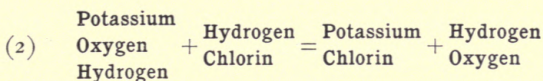
a. Try the effect of hydrochloric acid on the following metals that you have studied, — iron, magnesium, zinc, and copper, and verify the following statements: —



Name the salts formed.

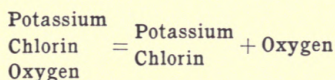
b. Neutralize a small amount of hydrochloric acid, first with sodium hydroxid, second with potassium hydroxid, and third with sodium carbonate, and verify the following statements: —





Name the salts formed.

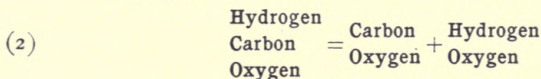
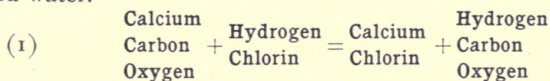
c. Potassium Chlorid. Compare the salt formed in *b 1* by neutralizing potassium hydroxid and hydrochloric acid, with the residue obtained in the hard glass tube in Experiment 6 *c*. We now see that potassium chlorate is composed of potassium, chlorin, and oxygen, and that the action in Experiment 6 *c*, can be expressed by the following:—



EXPERIMENT 26

Carbon Dioxid from Marble and Hydrochloric Acid

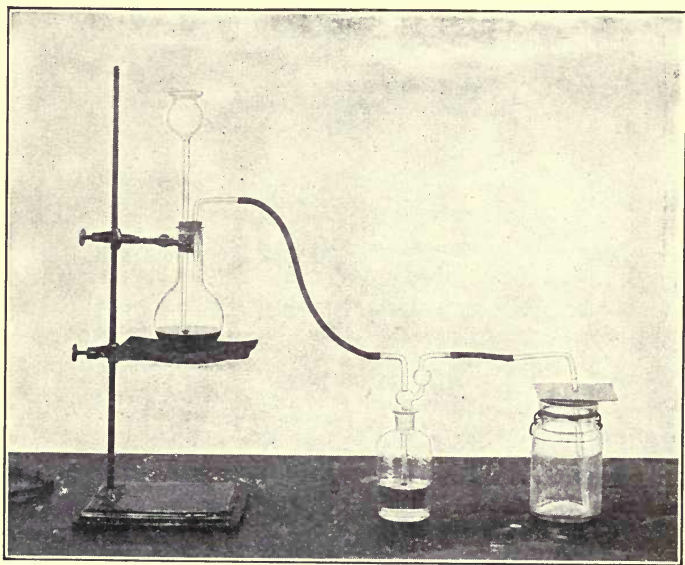
The best way to obtain a stream of carbon dioxid for use in the laboratory is by the action of hydrochloric acid upon marble (calcium carbonate). In a 500 cc. flask fitted with funnel and exit tubes, place a number of lumps of marble, and then add dilute hydrochloric acid (1 part acid to 1 part water). Carbonic acid is first formed, which, being very unstable (see Experiment 8 *c*), breaks up into carbon dioxid and water.



EXPERIMENT 27

Preparation of Chlorin

a. Test for Manganese. Take a piece of platinum wire, and make a loop on one end about 3 mm. in diameter. Heat it in the Bunsen flame, and, while it is still hot, dip it into a little microcosmic salt. Place it in the flame again, and make a bead of the fused salt. Melt into it a grain of powdered manganese dioxid, and, after holding in the flame for a few



moments, remove it, and note the characteristic amethyst color. Try other manganese compounds in the same manner.

b. Preparation. Under a hood, place a flask containing six or seven small lumps of manganese dioxid, and add

enough hydrochloric acid to cover them. Warm the flask, and recognize the gas evolved as chlorin. (See illustration on page 40.) This is the way your instructor prepared the gas for you. After as much gas as possible has been evolved, filter the liquid, evaporate, and crystallize. Put a crystal or two in a test tube, and add a little concentrated sulfuric acid. Notice that hydrochloric acid is evolved. Therefore, since hydrochloric acid contains chlorin, and, since that chlorin could not have come from the sulfuric acid, it must have come from the crystals. Make a micro-cosmic bead on a loop of platinum wire as in *a*, add a little of the crystals, and heat in the hottest part of the flame. The characteristic amethyst color proves the presence of manganese in the crystals. It is evident, therefore, that the crystals are manganese chlorid. The only way we can account for the oxygen of the manganese dioxid is that it united with the hydrogen of the hydrochloric acid and formed water. Therefore we have the following:—

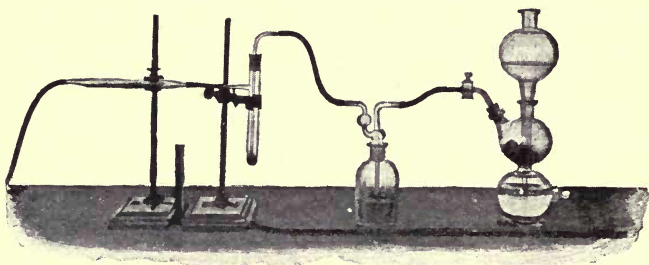


EXPERIMENT 28

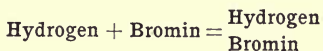
Bromin

a. Properties. Examine under a hood a small amount of bromin, and note its properties,—color, odor, weight, etc. Be careful not to inhale the fumes, since bromin acts violently upon the membranes of the throat and lungs. A drop of bromin on the skin produces a severe wound. Note its solubility in water, alcohol, ether, and carbon bisulfid. Have no fire near, since alcohol, ether, and carbon bisulfid are very inflammable. Hang moist pieces of litmus paper in vapors of bromin, and note the bleaching action.

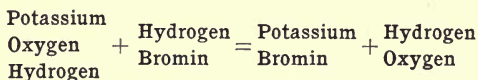
b. Hydrogen Bromid (Hydrobromic Acid). Have ready under a hood a test tube containing about 1 cc. of bromin. Fit this with a two-holed stopper containing an entrance and an exit tube. Let the entrance tube extend below the surface of the bromin. Attach the entrance tube to a hydrogen generator, so that the hydrogen must bubble through concentrated sulfuric acid in a catch bottle.



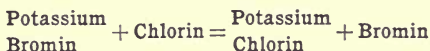
Attach the exit tube to a hard glass tube about 20 cm. long, containing platinized asbestos. Fit a rubber tube about 20 cm. long to this. Allow the hydrogen to pass through the apparatus until all the air is expelled. Then heat the platinized asbestos to redness, warming the bromin very gently. Collect the new gas in a jar by displacement of air, and get its properties. This gas is called hydrogen bromid, or hydrobromic acid.



c. Hydrobromic Acid and Potassium Hydroxid. Neutralize a solution of hydrobromic acid with potassium hydroxid, and verify the following: —



d. Replacement of Bromin by Chlorin. Dissolve about a gram of sodium or potassium bromid in water. After adding a little chlorin water and about 1 cc. of ether or carbon bisulfid, shake it vigorously. Which is the stronger chemically, chlorin or bromin?



e. Potassium Bromid and Sulfuric Acid. Try to make hydrobromic acid, using sulfuric acid and potassium bromid.

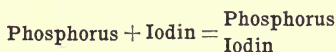
EXPERIMENT 29

Iodin

a. Properties. Take a few crystals of iodine, examine, and note properties. Warm a small crystal in a test tube, and note the color, odor, and specific gravity of the vapor. Test the solubility of iodine in water, alcohol, ether, and carbon bisulfid. Take a small lump of starch, add a few drops of water and mix into a paste. Then heat a test tube of water to boiling, and add this quickly to the starch. Place a little of the starch solution in a test tube of water, shake it, and add a little of the water solution of iodine. The blue color obtained is a characteristic test for the presence of uncombined iodine.

b. Phosphorus Iodid. Weigh out accurately about 1 gm. of phosphorus, and dissolve it in about 5 cc. of carbon bisulfid in a test tube. Then weigh out 8.2 times as much iodine. Add the iodine gradually to the phosphorus solution. When the iodine has disappeared, drive off the carbon bisulfid by placing the test tube in warm water. Do this

under a hood. The resulting crystals must be phosphorus iodid.*

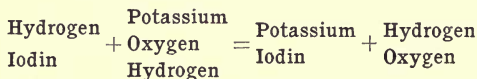


Hydriodic Acid. By passing vapors of iodine and hydrogen over finely divided platinum raised to a red heat, the two elements may be made to combine, but not readily. When made by this or other methods, hydrogen iodide is a colorless, heavy gas, very soluble in water, fumes in moist air, and gives an acid reaction with litmus paper. It is very similar to hydrochloric and hydrobromic acids.

EXPERIMENT 30

Potassium Iodide

a. Hydriodic Acid and Potassium Hydroxide. Neutralize a small quantity of dilute hydriodic acid with potassium hydroxide, and verify the following:—



b. Potassium Iodide and Sulfuric Acid. In a test tube, place a crystal of potassium iodide. Add a little concentrated sulfuric acid. Do you get hydriodic acid?

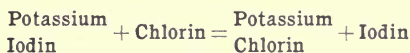
c. Potassium Iodide and Phosphoric Acid. In a test tube, place a crystal of potassium iodide, and add a little strong phosphoric acid. Heat it, and see whether you obtain hydriodic acid.

* When phosphorus iodide is thrown into water, hydriodic acid is formed, part of the hydrogen of the water uniting with the iodine of the iodide.

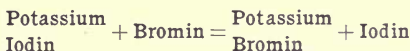
d. Potassium Iodid and Starch Paste. Try the effect of starch paste upon potassium iodid.

e. Solubility of Iodin in Potassium Iodid Solution. Dissolve a little potassium iodid in water in a test tube. Add a crystal of iodine. Is iodine soluble in potassium iodid?

f. Replacement of Iodin by Chlorin and Bromin. Dissolve a little potassium iodid in water in a test tube. Add chlorine water, and separate into two equal parts. To one add starch paste, to the other add carbon bisulfid, and shake them. Which is the stronger chemically, chlorine or iodine?



Try the same, using bromine water instead of chlorine water. Which is the stronger chemically, bromine or iodine?



Fluorine.

There is an element fluorine, a colorless gas, which in a great many ways is similar to chlorine, bromine, and iodine. It is the most active chemically of all the elements, and was not known in the uncombined state until comparatively recently. One great difficulty in preparing it was to find material for vessels that it would not attack. Its principal salt is calcium fluoride, which occurs in nature as the mineral fluor spar.

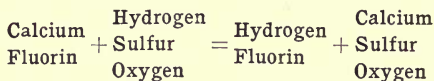
EXPERIMENT 31

Calcium Fluoride and Hydrogen Fluoride

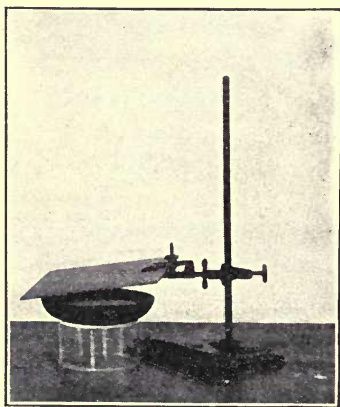
a. Properties of Calcium Fluoride. Examine calcium fluoride both in crystal and powder form.

b. Hydrogen Fluoride (Hydrofluoric Acid). In a test tube,

put 1 cc. of powdered calcium fluorid. Cover with sulfuric acid, and heat it under the hood. Get the color and odor of the gas, being careful to smell cautiously, since the gas seriously affects the membranes of the throat and lungs. Breathe over the mouth of the test tube. Test the gas with blue litmus paper. Remove the contents of the tube, wash and dry it. Examine the glass, noticing how it has been corroded. This gas is called hydrofluoric acid. Assuming that calcium fluorid is composed of calcium and fluorin, we readily see that



c. Etching Glass. In a shallow lead dish, make a paste of powdered calcium fluorid and concentrated sulfuric acid.



Take about half as many cc. of acid as you take grms. of calcium fluorid. Cover a piece of window-glass with a thin layer of melted paraffin, warming the glass to spread the paraffin evenly. When the glass is cool, trace a figure on it, and, laying it upon the dish, leave it over night.

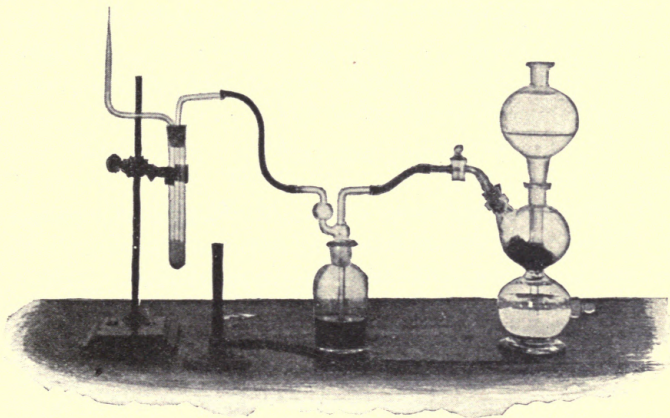
Glass is composed of various substances, among which is the element silicon.

Fluorin has a strong affinity for this element; hence it decomposes the glass by abstracting the silicon, forming a colorless gas, silicon fluorid. Thus the hydrofluoric acid corrodes the glass where it was not protected by the paraffin.

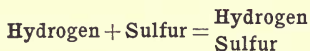
EXPERIMENT 32

Sulfids

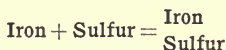
a. Synthesis of Hydrogen Sulfid. Have ready a hydrogen generator. Take a test tube, place in it about 10 grms. of sulfur, and fit it with a two-holed stopper containing an entrance tube leading within 2 cm. of the sulfur, and an exit tube bent at right angles twice, the second bend extending



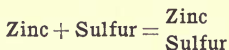
upwards about 10 cm. and drawn to an opening of about 1 mm. diameter. Connect with the hydrogen generator so that the hydrogen must pass through a catch bottle containing concentrated sulfuric acid, to dry it. Allow the hydrogen to pass through the apparatus in order to drive out the air; then, when all is safe, heat the sulfur to boiling. Note the odor of the escaping gas. Light it, and get the color of the flame. Heat the tube gently above the stopper, and note that sulfur is deposited. Is hydrogen strongly united with sulfur in hydrogen sulfid?



b. Sulfids of Metals. 1. Mix equal volumes of fine iron filings and flowers of sulfur. Heat the mixture in a test tube. After action has ceased, break open the tube and get the properties of the resulting iron sulfid.

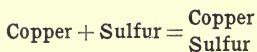


2. Mix 6.5 parts of zinc dust and 3.2 parts of the flowers of sulfur, and place the mixture on an iron plate. In mixing, do not use pressure, since the two elements may unite with explosive violence. *Cautiously* light the mixture, and get the properties of the resulting zinc sulfid.



3. Place in a test tube a mixture of 6.4 parts of copper filings with 3.2 parts of flowers of sulfur, and heat as in 1. Get the properties of the resulting copper sulfid.

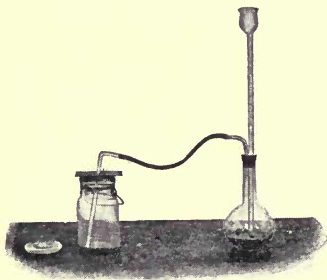
In a test tube, boil a small quantity of sulfur. By means of a pair of long forceps, hold a strip of copper in a flame until red hot; then quickly hold it in the vapors of sulfur in the tube.



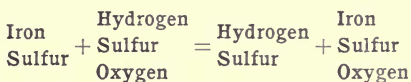
How do these *sulfids* differ chemically from the corresponding *sulfates*?

c. Hydrogen Sulfid from Iron Sulfid and Sulfuric Acid.

In a 100 cc. flask, place about 10 grms. of iron sulfid in small lumps. Add dilute sulfuric acid (3 parts acid to 1 part

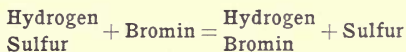


water), and warm it. Collect in jars and recognize the same gas as was obtained in *a*, i.e., hydrogen sulfid. Is it soluble in water? Note its action on blue litmus paper. Ignite a jar of the gas, and note what the products of the combustion are. When all the gas possible has been evolved, filter, evaporate, and crystallize it. Recognize the crystals as *iron sulfate*. (See Experiment 15 *f*.)



If you had used the sulfids made in 2 and 3, what salts would you have obtained instead of iron sulfate?

d. Action of Bromin upon Hydrogen Sulfid. In a 100 cc. flask, place about 50 cc. of water; add a drop of bromin, and shake it until it is dissolved. Fit the flask with an exit and entrance tube, and, under a hood, pass hydrogen sulfid through the liquid. Note the disappearance of the red color. Filter the solution; then recognize the residue as sulfur by holding it with a pair of forceps in a Bunsen flame.

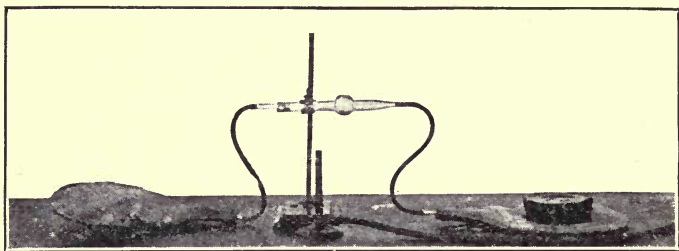


EXPERIMENT 33

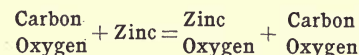
Other Compounds of Carbon

a. Carbon Monoxid. Fill a rubber gas-bag with dry carbon dioxid made from marble and hydrochloric acid (see Experiment 26). Place in a hard glass combustion tube about 5 grms. of zinc dust, and clamp this to a support. To one end, attach an empty bag similar to the first; to the other end, the bag filled with the gas. Heat the tube with a blast lamp, and pass the carbon dioxid back and forth over the hot

zinc. Notice the color of the compound forming in the tube, both when it is hot and cold. It evidently corresponds to the zinc oxid formed in Experiment 12 *b*. The only substance from which the zinc could have obtained so much oxygen must be the carbon dioxid. Has it removed all the oxygen from the gas? If so, there would be nothing but carbon left. Force all the gas into one bag and confine it by means of a

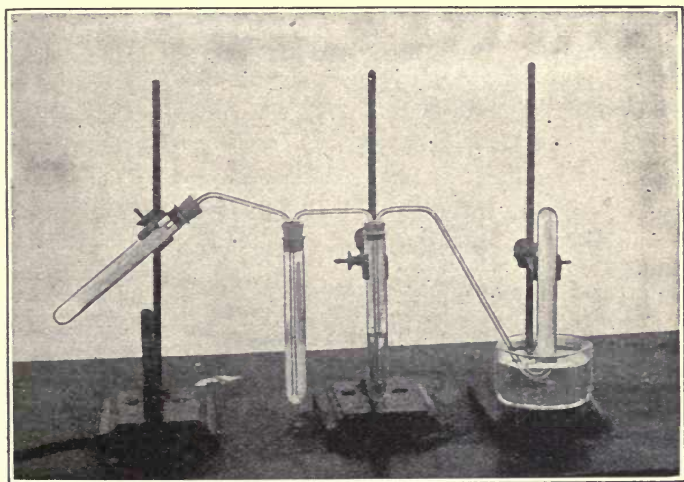


pinch cock. In a glass vessel filled with a moderately strong solution of sodium hydroxid, invert a beaker filled with the liquid, and collect some of the gas from the bag. The solution of sodium hydroxid will absorb any carbon dioxid remaining unchanged (see Experiment 19 *a*), and will allow the new product to pass through. Get the properties of the new gas. Lift the beaker from the solution; immediately ignite the gas with a match, and note the color of the flame. This new gas must be composed of carbon and oxygen, but it must have less oxygen than carbon dioxid, since the zinc removed a part. It is called *carbon monoxid*. Which substance is reduced? Which is oxidized?

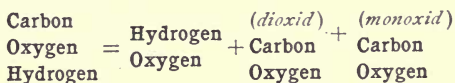


b. Analysis of Oxalic Acid. In a hard glass tube closed at one end, and fitted with a one-holed stopper containing an exit tube, place about 5 grms. of oxalic acid crystals. Con-

nect the exit tube with a dry empty test tube placed in cold water, and with this connect a third test tube containing lime water, so that any gas coming from the hard glass tube must pass to the bottom of each test tube. Now heat the hard glass tube very gently, having it inclined at an angle of 30° , and collect the evolved gas over a sodium hydrate solution, as in *a*. Recognize this gas as carbon monoxid. Note the formation of water in the first test tube. Note also the milki-ness which is at first produced in the lime water, but which



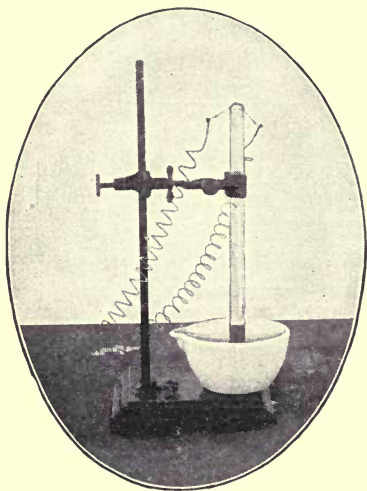
soon disappears, proving the presence of carbon dioxid. The sodium hydrate removes all traces of this from the carbon monoxid. Therefore, we see, oxalic acid is composed of water, carbon dioxid, and carbon monoxid, i.e., carbon, oxygen, and hydrogen.



c. Carbon Monoxid from Oxalic Acid. In a 250 cc. flask fitted with a funnel and exit tube, place about 20 grms. of oxalic acid crystals. Add through the funnel tube enough concentrated sulfuric acid to cover the crystals. Heat it, and pass the evolved gas first through an empty catch bottle, and then through a second containing a strong solution of sodium hydrate. Collect the gas over water, and recognize as carbon monoxid.

We know that sulfuric acid has a strong attraction for water. If, then, it takes water away from oxalic acid, we shall have only the two gases carbon dioxid and carbon monoxid left. The carbon dioxid is absorbed by the sodium hydroxid in the catch bottle.

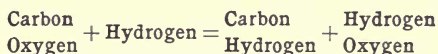
d. Methyl Hydrid (Marsh Gas) from Carbon Monoxid and Hydrogen. In a eudiometer tube fitted with two platinum



wires fused into the closed end, collect over mercury 10 cc. of carbon monoxid and 30 cc. of hydrogen, both gases being dry. Allow electric sparks from an induction coil to pass between the points. Notice the formation of moisture and the decrease in volume. Get the properties of the new gas. It is called methyl hydrid, or marsh gas, and is composed of carbon and hydrogen. Part of the hy-

drogen must have united with oxygen to form water, and part must have united with the carbon to make the new

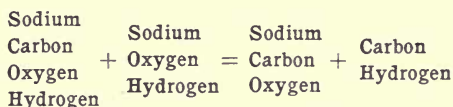
gas; otherwise only water and carbon would have been formed.



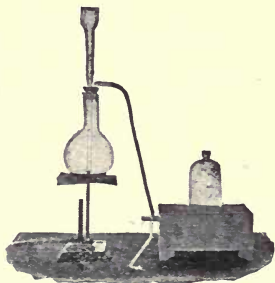
e. Marsh Gas from Sodium Acetate and Soda Lime. In a hard glass tube, place an intimate mixture of one part sodium acetate and four parts soda lime. Heat it until gas begins to form, and then keep the temperature as constant as possible. Collect the gas over water, and recognize it as the same gas as in *d*, i.e., marsh gas. Dissolve the contents of the tube when cool in water. Filter, crystallize, and recognize as sodium carbonate. The calcium oxid in the soda lime is added only to prevent fusion.

It is difficult to show to the beginner the composition of sodium acetate. Suffice it to say however that it is composed of sodium, carbon, oxygen, and hydrogen.

Hence we have

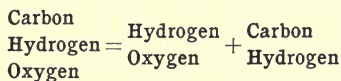


f. Ethylene. In a 1000 cc. flask fitted with a funnel and exit tube, place 10 cc. of 95% alcohol. Add 55 cc. of concentrated sulfuric acid and heat gently. Collect the gas evolved over water and get its properties. Have ready a eudiometer tube as in *d*, and fill it one fourth full of the gas dried by being passed through sulfuric acid. Allow the sparks from an induction coil to pass through the gas, and notice the carbon deposited upon the platinum



points. Recognize the remaining gas as hydrogen. This gas is called ethylene, or olefiant gas.

Since sulfuric acid has a strong affinity for water, we are justified in believing that the action of the acid upon the alcohol was to abstract hydrogen and oxygen from it in the form of water. Hence alcohol is composed of carbon, oxygen, and hydrogen.



Remark. Carbon and hydrogen unite in different proportions, and form a large class of compounds called hydrocarbons. These are more properly studied in that division of chemistry called Organic Chemistry.

EXPERIMENT 34

Nature of Flame

a. Light a candle, and press a glass plate down upon the flame. Notice that combustion is taking place only on the outside of the flame. Remove the glass plate, and, into the dark interior, insert the end of a small tube about 1 mm. bore. Hold a lighted match to the other end, and notice the flame. The dark interior must be composed of unburned gas. Note the unburned carbon (soot) on the glass plate. Does white-hot carbon give light? To what is the luminosity of the flame due?

b. Examine the flame of a Bunsen burner in the same way when the air holes are shut. Draw a diagram of the Bunsen flame when the air holes are open.

c. Hold an inverted test tube over a candle or Bunsen flame for a few moments. Remove the tube, and test the

gas in it with lime water, proving the presence of carbon dioxid. Hold a cold glass plate over a flame for a moment, and note the condensed water. Hence two products of the burning are carbon dioxid and water.

d. Bring down a piece of wire gauze upon the flame of a Bunsen burner. Note that there is no flame above the gauze. Now hold a lighted match over the gauze. Was unburned gas passing through the gauze? Now turn off the gas altogether. Hold the gauze about two inches above the burner. After turning on the gas again, hold a lighted match above the gauze. The gas will now burn above the gauze, but not below it.* Perform the first part of this experiment a second time, and hold the gauze on the flame for some time. After a while the gauze becomes hot, and the gas above it takes fire.

e. Have ready a platform balance. Cut a piece of wire netting ten inches long by five inches wide. Bend the netting into the form of a cylinder, so that its longer side is the altitude of a cylinder. Divide the cylinder into two equal parts by means of a circular piece of netting fastened inside the cylinder by means of wire. Fill the upper half about one quarter full of broken pieces of sodium hydroxid. Then place upon this a second circular piece of netting, and, in the compartment thus made, place an equal quantity of fused calcium chlorid. Place the cylinder upon one platform of the balance with the empty half over a candle about

* The mine's safety lamp depends upon this principle. This lamp is simply an ordinary oil lamp having the flame surrounded by wire gauze. The explosive mixtures of gases (air and compounds of carbon and hydrogen), which occur in mines, may pass through the gauze and burn inside the lamp; but outside, on account of the relatively high "kindling temperatures" of carbon and hydrogen, they do not take fire.

three quarters of an inch long, and, upon the other, place weights enough to balance it exactly. Light the candle, and watch the effect upon the equilibrium. Be careful not to place the apparatus in a draught. What is the relation between the weight of the products formed by the burning candle and the weight lost by the candle?

EXPERIMENT 35

Hard and Soft Water

a. Permanently Hard Water. Place about 5 grms. of calcium sulfate in 50 cc. of distilled water in a beaker; stir, and filter it. Place 10 cc. of this water in a test tube, and add from a soap solution,* drop by drop, until a permanent froth is obtained, counting the number of drops necessary. In the same way, test 10 cc. of distilled water. Now boil the remaining 40 cc. in the beaker, and, when it is cool, test 10 cc. of it again with soap solution. Does boiling make the water soft?

b. Temporarily Hard Water. Into a beaker containing 50 cc. of lime water, pass carbon dioxide until the precipitate first formed is dissolved. Test 10 cc. of this with soap solution the same as in *a*. Boil the remaining 40 cc. until the precipitate all returns. Then filter and test 10 cc. of it with soap solution. Has the water become soft?

Remark. Water containing calcium sulfate in solution is called *permanently hard* water. It cannot be made soft by boiling, since calcium sulfate is about as soluble in hot water as in cold.

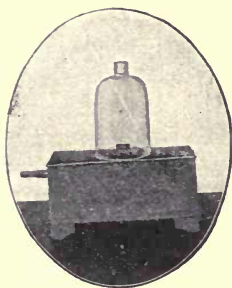
* To make a soap solution, place 1 grm. of castile soap shavings in 10 cc. of distilled water, and dissolve as much as possible. Then pour off the clear solution.

Water containing calcium carbonate held in solution by carbonic acid is called *temporarily hard* water, since it can be made soft by boiling, thus driving off the carbon dioxid and precipitating the calcium carbonate.

EXPERIMENT 36

Nitrogen

Float on water in a pneumatic trough a flat cork fitted with the bowl of a deflagrating spoon containing a piece of phosphorus about the size of a pea. Ignite the phosphorus, and hold a bell jar over it. Place a glass plate under the jar, when it is cool; keeping the water inside and outside at the same level; remove the jar, and shake it. Note that one-fifth of the air was oxygen. Get the properties of the remaining four-fifths of the air. This gas is called nitrogen.



Remark 1. Air is a mechanical mixture and not a chemical compound. The following are the best proofs of this statement:—

1. Nitrogen and oxygen are not present in the proportions in which they would be, if they were chemically united.
2. The composition of air is slightly variable. (See Experiment 2, Part II.)
3. A mixture of nitrogen and oxygen in the proportion of 4 to 1 has the same properties as air.
4. Air is slightly soluble in water. But air dissolved in water does not contain the same proportions of nitrogen and oxygen as atmospheric air. The water takes up the two gases according to their respective solubilities.

Remark 2. Since organized bodies, both vegetable and animal, all have as their principal constituent the element carbon, the oxidation of these bodies, either by the rapid process of combustion or the slower one of decay, must be accompanied by an evolution of carbon dioxid. From these sources the carbon dioxid in the atmosphere is obtained. When the life process, whatever it may be, ceases to act, the oxygen of the air, aided by *micro-organisms* known under the different names of *germs*, *microbes*, *bacilli*, etc., immediately attacks the compounds of which the body is composed, and, by more or less complicated processes, changes them into other compounds, one of which is carbon dioxid. Even during life the process goes on, and sometimes it is absolutely necessary to the living organism, as in the case of the oxidation that is taking place continually in the bodies of animals. The oxygen of the air is taken into the lungs, where it is taken up by the blood. A certain substance called *haemoglobin*, which is found in the red blood corpuscles, unites with the oxygen, and thus it is carried to the remotest part of the body. The haemoglobin gives up its oxygen wherever there is material whose oxidation the economy of the life process requires. The products of this oxidation are carried back to the lungs, where they are expelled into the air. The chemical energy, stored in the compounds that are thus oxidized, is transformed into heat, and, in this manner, keeps up the necessary warmth of the body. The carbon dioxid that is expelled into the air from the lungs of living animals would soon increase the quantity of that gas in the atmosphere, were it not for the reciprocal action of plants. Just as animals require oxygen, so plants require carbon dioxid. The leaves of the plant admit the carbon dioxid through small apertures on their under sides. In the leaves there is a green substance called *chlorophyll*. This,

acted upon by sunlight, decomposes the carbon dioxid; and the carbon, together with hydrogen, oxygen, and other elements that have been obtained through the roots in the form of water and other compounds, is deposited throughout the plant structure. The oxygen that is left from the carbon dioxid is expelled into the air. In this manner the relative quantities of oxygen and carbon dioxid in the air are kept practically constant.

It is well to note here the difference between *organic* bodies and *organized* bodies. The life process shapes the *organic* matter into structures called *cells*, which, taken together, make up the *organized* body. The study of the chemistry of plant and animal life belongs to a special branch of chemistry, called Physiological Chemistry.

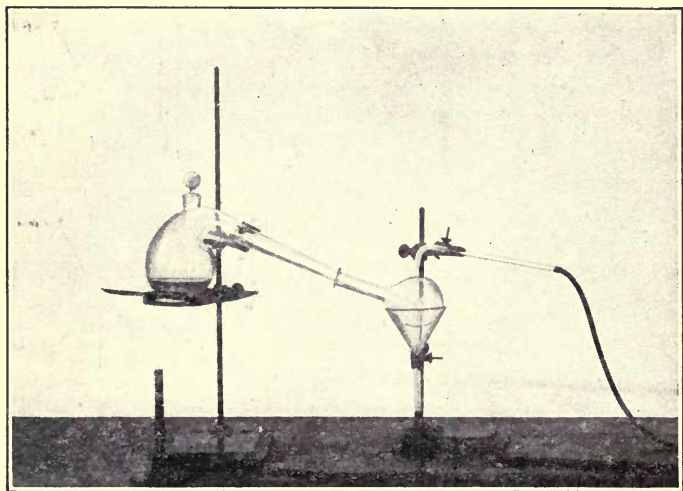
EXPERIMENT 37

Nitric Acid

a. Preparation. In a glass-stoppered retort, place 30 grms. of powdered sodium nitrate (Chili saltpeter). Add 10 cc. of concentrated sulfuric acid, being careful to allow none of it to fall into the neck of the retort. Have the neck of the retort extending into a 100 cc. flask over which cold water is kept running. Heat the retort gently, and distill over the liquid. Get its properties. Dilute a little with water, and test with blue litmus. This is called *nitric acid*. When the residue is cool, dilute it in the retort with water, evaporate, crystallize, and recognize it as sodium sulfate.

b. 1. Analysis of Nitric Acid. Pour into a 100 cc. flask, fitted with a one-holed cork stopper and delivery tube, a little dilute nitric acid, and add a few pieces of magnesium ribbon. Catch the gas evolved, and recognize it as hydrogen.

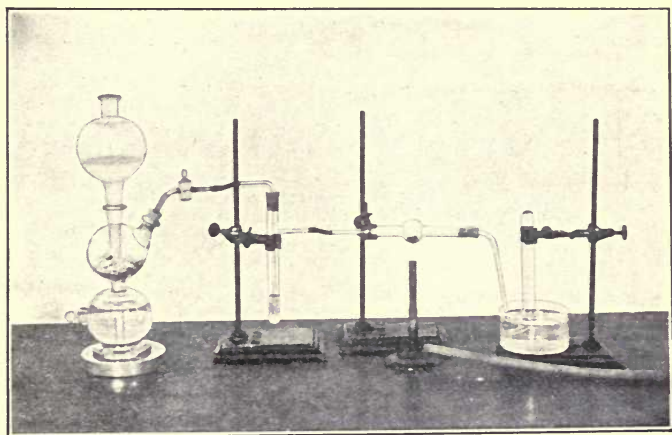
Since the hydrogen could not have come from the magnesium, or from the magnesium and water, it must have come from the nitric acid. Hence one constituent of nitric acid is hydrogen (see def. acid, page 29).



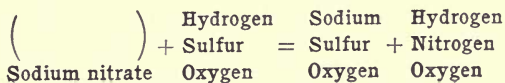
2. Have ready an apparatus for obtaining a steady stream of dry carbon dioxide (see Experiment 26). Connect this with a test tube fitted with an exit tube and an entrance tube extending almost to the bottom. Place about 4 cc. of concentrated nitric acid in the test tube. Connect the test tube with a combustion tube containing clean copper filings. Have ready a glass vessel containing a moderately strong sodium hydroxid solution. Pass carbon dioxide through the apparatus to remove the air; then gently heat the nitric acid, having meanwhile heated the copper filings to redness. Collect the gas evolved over the sodium hydrate solution, and

recognize it as nitrogen. The sodium hydroxid absorbs the carbon dioxid, which is simply used as a carrier for the nitric acid fumes. Note the formation of copper oxid in the combustion tube.

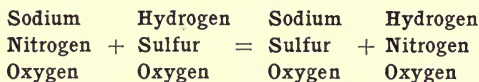
Since copper does not affect carbon dioxid, the oxygen must have come from the nitric acid. Therefore nitric acid contains nitrogen and oxygen. We then have as the component parts of nitric acid, hydrogen, nitrogen, and oxygen.



Argument. In the making of nitric acid we started with the sodium nitrate, a substance whose composition we do not know. This acted with sulfuric acid, and a compound containing sodium, sulfur, and oxygen, i.e., sodium sulfate, was formed, together with another compound containing hydrogen, nitrogen, and oxygen. We then have —



Knowing that the tendency of metals is to replace the hydrogen of acids, we readily see that the action was



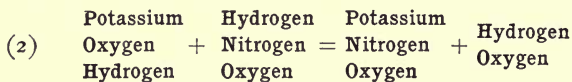
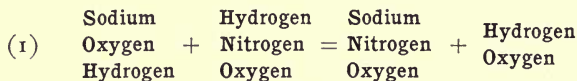
If this be true, we ought to obtain sodium nitrate by neutralizing nitric acid with sodium hydroxid. (See Experiment 38.)

Remark. Nitric acid, when left exposed to the air, gives off colorless fumes. When heated or exposed to sunlight, it decomposes slowly into water, oxygen, and various oxids of nitrogen. This is the reason the brown fumes are seen above the liquid in nitric acid bottles. Nitric acid is a very valuable oxidizing agent on account of the ease with which it gives up its oxygen. The compounds of nitric acid with organic substances are especially unstable.

EXPERIMENT 38

Neutralization of Nitric Acid

Neutralize nitric acid with sodium and potassium hydroxids.

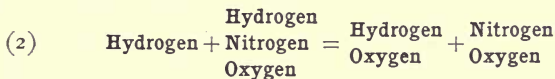
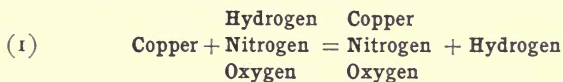


EXPERIMENT 39

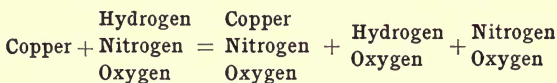
Nitric Oxid

a. Preparation. Place about 50 grms. of sheet copper clippings or else copper turnings in a 250 cc. flask fitted with funnel and exit tubes. Pour in through the funnel tube just enough nitric acid, diluted with an equal volume of water, to cover the copper. Collect the colorless gas in jars over water, and note its properties. Open a jar in the air, and note the properties of the new brown gas thus formed. Ignite a piece of yellow phosphorus about the size of a bean in a deflagrating spoon, and hang it in a second jar of the gas as in Experiment 7 *b*. When the burning has ceased and the jar is cool, and after the phosphoric oxid has settled, recognize the remaining gas as nitrogen (see Experiment 36). Therefore the two gases, the colorless and the brown, must be oxids of nitrogen, the latter containing the more oxygen.

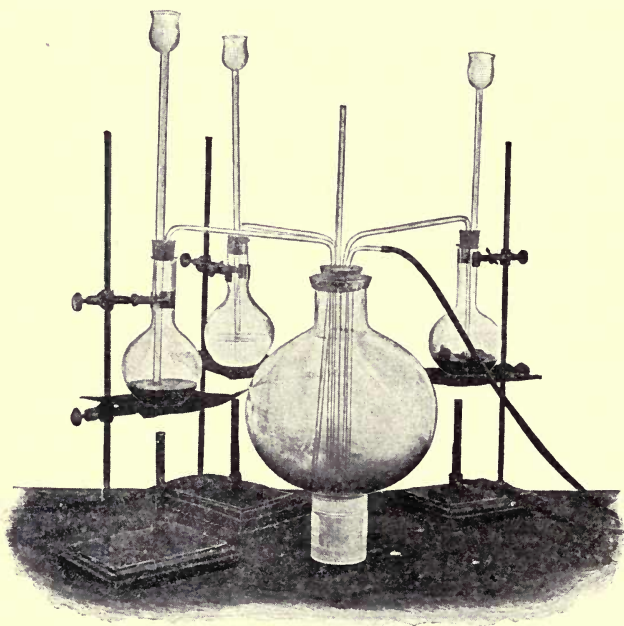
Here we have a case where apparently a metal does not replace the hydrogen of the acid. The fact of the matter is that elements at the moment they are released from their compounds will unite with other elements, or decompose compounds that they would not affect at other times. In this case, the hydrogen we should naturally expect to be evolved attacks the extra nitric acid instead, forming water and nitric oxid. We may express the two actions thus:—



Or we might write the two together thus :—



b. Sulfuric Acid by means of Nitric Oxid. Have ready a large flask (capacity about 3 or 4 liters), loosely fitted with a five-holed stopper. Let four of the holes contain entrance tubes, bent at right angles, and extending to the bottom of



the flask. Let the fifth hole contain an exit tube. Fit three 250 cc. flasks with funnel tubes and exit tubes, and arrange them so that they may be connected with three of the entrance tubes of the large flask. Fill the first of the small

flasks about half full of water; in the second, place copper and sulfuric acid (as in Experiment 9 *g*); and in the third, place copper and nitric acid (as in *a* of this experiment). Connect the fourth entrance tube with a pair of bellows. Allow steam, air, nitric oxid, and sulfur dioxid to enter the large flask. Note the decolorization of the brown gas that was formed when the nitric oxid first entered. Evidently the oxygen that the nitric oxid took on has been taken away by some other compound. Steam does not have that property; hence it must have been the sulfurous acid formed by the sulfur dioxid and water (see Experiment 9 *d*). But if sulfurous acid takes on oxygen, it becomes sulfuric acid (see Experiment 9, *d* and *f*).

After the operation has been going on for some little time in the large flask, remove the liquid that has collected, and recognize it as sulfuric acid.

This experiment illustrates the principle that is used in the manufacture of sulfuric acid on a large scale. (See page 173).

EXPERIMENT 40

Elements in Nascent State

Make a solution of a few crystals of potassium permanganate. Through this, pass a stream of hydrogen from a hydrogen generator. Is there any change in color? Then remove the generator, and add a little concentrated sulfuric acid. Is there now any change in color? Add to this a few pieces of zinc. What results?

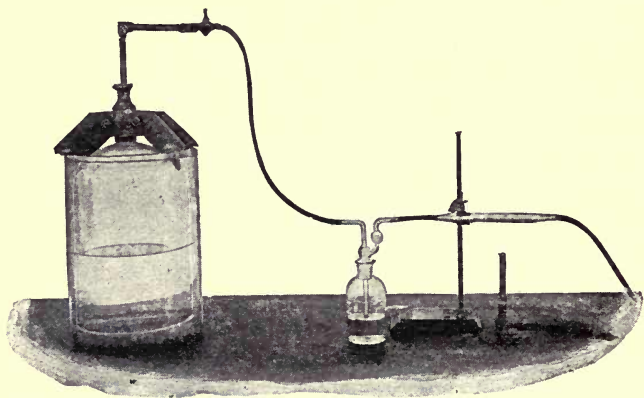
Here the hydrogen that first passed through the solution evidently had no effect. In the second case, where the hydrogen was set free by the zinc in the presence of the per-

manganate, it evidently decomposed it. Elements at the instant they leave their compounds have a stronger chemical attraction for other elements, and are said to be in the *nascent state*.

EXPERIMENT 41

Ammonia

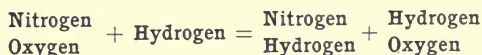
a. Ammonia from Nitric Oxid and Hydrogen. Fill a gas holder with two volumes of nitric oxid and five volumes of hydrogen, made respectively from copper and nitric acid, and zinc and sulfuric acid. Place loosely in a glass com-



bustion tube about 2 grms. of platinized asbestos, and connect the tube with the gas holder in such away that the mixed gases must pass through sulfuric acid in a catch bottle. To the other end of the tube, attach a piece of rubber tubing about two feet long. Allow the mixed gases to pass slowly through the apparatus until the air is expelled; then, when no brown color is seen, heat the platinized asbestos with a

Bunsen flame. Notice the water that collects on the cooler part of the tube. Get the odor of the gas that escapes, and test the gas with moist red litmus paper.

Since, in the mixed gases, there are only the elements hydrogen, nitrogen, and oxygen, and since water is formed, the escaping gas must be either nitrogen, hydrogen, oxygen, a compound of nitrogen and oxygen, or else a compound of nitrogen and hydrogen. On account of its properties, it cannot be any of the first four (see Experiments 36, 15 a, 6, and 39), so it must be composed of nitrogen and hydrogen. We call the gas ammonia.



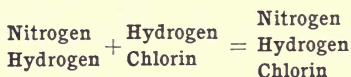
b. Ammonia from Organic Matter containing Nitrogen. Make a mixture of freshly slacked lime and dry sodium hydroxid. Cut up a piece of flannel or horn into small bits, and heat the pieces in a test tube with the above mixture. Recognize the gas that escapes as ammonia.

EXPERIMENT 42

Ammonium Chlorid

Allow ammonia gas and hydrochloric acid gas to come in contact with each other, and note the white fumes that condense in solid form. Taste the new substance.

This new compound must be composed of nitrogen, hydrogen, and chlorin; but there must be more hydrogen in it than in ammonia. To indicate this we name the substance ammonium chlorid.

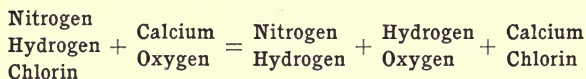


EXPERIMENT 43

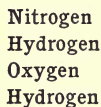
Ammonia from Ammonium Chlorid

Mix intimately 20 grms. of ammonium chlorid and 40 grms. of calcium oxid. Put the mixture into a copper retort,* or into a large piece of a gas pipe closed with a plug at one end. Connect this with a tube containing calcium chlorid to dry the evolved gas, and after heating recognize the gas as ammonia. Note that the calcium chlorid absorbs water. Collect several jars full by displacement of air, i.e., allow a delivery tube to reach to the bottom of the jar, but have the jar inverted, since ammonia is lighter than air. Open a jar under water, and note the great solubility of the gas.

Since water and ammonia are formed by the chemical action, the only elements left to account for are the calcium and the chlorin. These must have united to form calcium chlorid, which action may be verified by dissolving the contents of the retort, filtering and crystallizing. Therefore we have



Remark. When ammonia dissolves in water, we call the product ammonium hydroxid, that is,



* If the chlorid and lime are dry, the experiment may be performed in an ordinary flask.

The two elements nitrogen and hydrogen, as they appear in ammonium chlorid and ammonium hydroxid, are capable of passing from compound to compound without separating, and act just as if they were a metal. Such a combination of elements is called a *compound radical*.

EXPERIMENT 44

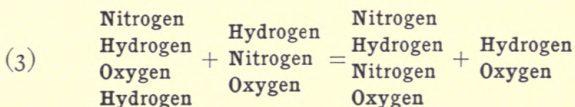
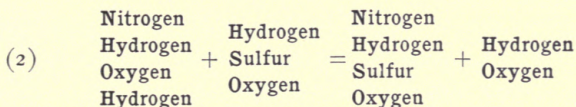
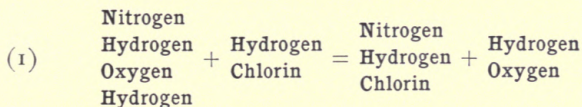
Ammonium Amalgam

To show how much like a metal the radical ammonium acts, it is possible to make an ammonium amalgam with mercury. Into a strong solution of ammonium chlorid, drop a piece of sodium amalgam (see Experiment 10 *d*). As the ammonium amalgam decomposes, notice the fumes of ammonia.

EXPERIMENT 45

Neutralization of Acids with Ammonium Hydroxid

Neutralize dilute hydrochloric, sulfuric, and nitric acids with ammonium hydroxid, and verify the following:—

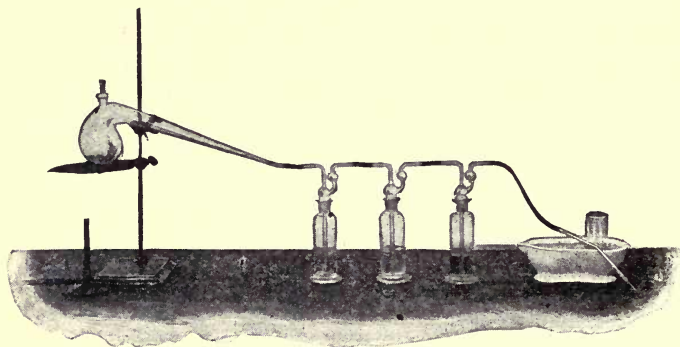


Evaporate the solution. Place a little of each of the salts formed in small tubes closed at one end, and heat. The phenomenon in the case of the first two is called *sublimation*. Name the salts.

EXPERIMENT 46

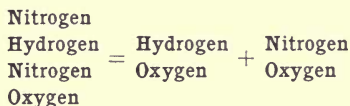
Nitrous Oxid

In a stoppered glass retort, place 15 grms. of ammonium nitrate. Connect the retort with three catch bottles, the first empty, the second containing a dilute solution of sodium hydroxid, and the third containing a solution of iron sulfate. Heat the retort gently. Note the formation of water in the



first catch bottle. The sodium hydroxid will absorb any chlorin which may be obtained from the commercial ammonium nitrate, and the iron sulfate will absorb any nitric oxid which will be formed if the ammonium nitrate is heated too hot. Collect the gas over warm water, since it is soluble in cold. Note its properties, especially odor, taste, and relation to combustion.

Since ammonium nitrate is composed of nitrogen, hydrogen, and oxygen, and since water is formed in the experiment, the new gas must be either nitrogen, a compound of nitrogen and hydrogen, or a compound of nitrogen and oxygen. On account of the properties of the first and second, and on account of its supporting combustion, it must be the third. Here we have another oxid of nitrogen, which is called nitrous oxid to distinguish it from nitric oxid and nitric peroxid.

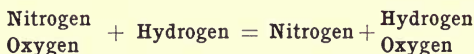


Remark. This oxid of nitrogen is commonly called laughing gas. When taken into the lungs, it acts as an anaesthetic, and is hence valuable in certain surgical operations. For this reason it is used by dentists. Its comparative instability is shown by the fact that it gives up its oxygen easily, thus supporting combustion. Whatever the forces are that hold elements together, they are evidently in this case in "*unstable equilibrium*," i.e., if a substance easily oxidized is brought in contact with nitrous oxid under proper conditions, the oxid gives up its oxygen. In the case of nitric oxid, however, the equilibrium seems to be more stable; and, although there is more oxygen in the compound, it is held there more firmly. From Experiment 39, we learn that nitric oxid will unite with more oxygen, thus giving a still higher oxid. It is therefore a reducing agent. On the other hand, the oxygen thus obtained by the nitric oxid is not held firmly, and the peroxid formed will give it up readily. Nitrogen peroxid is therefore an oxidizing agent.

EXPERIMENT 47

Analysis of Nitrous Oxid

Collect 5 cc. of the gas over mercury in a eudiometer tube with platinum wires fused in the closed end. Add to this an equal volume of hydrogen. Let both gases be dry. Pass sparks through the mixture by means of an induction coil. Recognize the remaining gas as nitrogen.



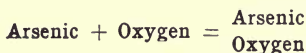
EXPERIMENT 48

Arsenic

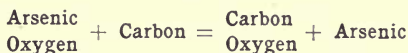
Perform all experiments with arsenic under a good hood. Be careful not to breathe any arsenic fumes, since they are *deadly poisons*.

a. Properties. Take a bit of the element arsenic, and note as many properties as you can. Place a small piece about 2 mm. in diameter in a glass tube closed at one end. Heat the tube in the Bunsen flame, noticing that arsenic passes immediately from the solid to the gaseous state, after which it again condenses on the cooler part of the tube, forming the so-called arsenic mirror.

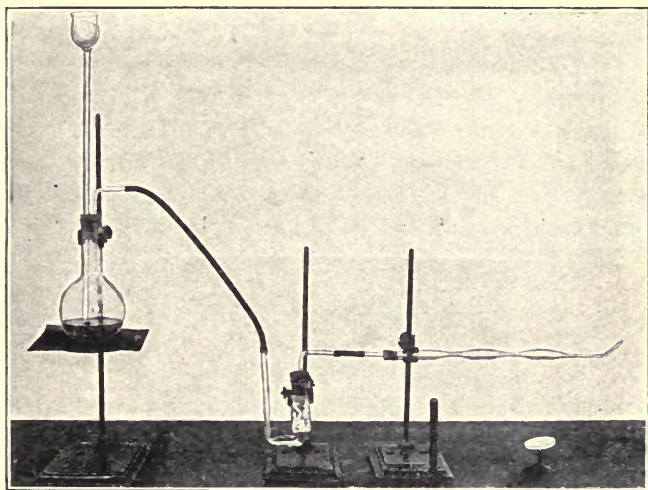
b. Arsenic Oxid. Oxidize a bit of arsenic about the size of a pin head under a hood. Be careful not to breathe the fumes. You can hardly avoid noticing the garlic-like odor of arsenic.



c. Reduction of Arsenic Oxid by means of Carbon. Mix a little arsenic oxid with powdered charcoal, and place it in a tube like that used in *a*. Heat the tube and verify the following:—



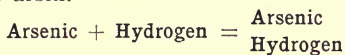
d. Arsin (Hydrogen Arsenid). Connect a flask for generating hydrogen (zinc and sulfuric acid) with a U tube con-



taining calcium chlorid, and then join this with a glass tube narrowed in several places and finally drawn to a small opening. When the air is expelled from the apparatus, light the hydrogen jet and hold a clean crucible cover in the flame for a moment. Then add a little arsenic or arsenic oxid through the funnel tube of the hydrogen generator. Note the change of color in the flame and the vapors of arsenic

oxid that arise.* Now again place the porcelain crucible cover in the flame, and note the arsenic mirror formed. Apply the flame of a Bunsen burner to one of the bulbs of the tube, and note the formation of the mirror in the narrow part, just beyond the heated portion. Test the mirror on the crucible cover with a solution of bleaching powder.

We have seen before that elements, when in the nascent state, combine more easily with other elements. Here we have nascent hydrogen uniting with arsenic, forming hydrogen arsenid, or arsin.



Remark. There are two acids of arsenic, similar in their chemical properties to phosphorous and phosphoric acids. They are called arsenious and arsenic acids. Both are composed of hydrogen, arsenic, and oxygen, the latter containing the greater amount of oxygen.

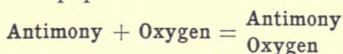
EXPERIMENT 40

Antimony

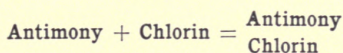
a. Properties. Examine a piece of antimony, and note its properties. See whether you can obtain a mirror, as with arsenic.

* In this experiment we have both *complete* and *incomplete* combustion of arsin. In the first case, the products of combustion are evidently water and arsenic oxid; in the second they are water, arsenic oxid, and arsenic. In the same way, when carbon bisulfid burns, we obtain either carbon dioxid and sulfur dioxid, or else carbon dioxid, sulfur dioxid, and sulfur; when hydrogen sulfid burns, we obtain either water and sulfur dioxid, or else water, sulfur dioxid, and sulfur; and when coal gas burns, we obtain either water and carbon dioxid, or else water, carbon dioxid, and carbon. It is evident, therefore, that the affinity of oxygen for these elements in combination is not so great as when they are free.

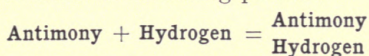
b. Antimony Oxid. Oxidize a small piece of antimony, and note the color of the oxid. Drop a bit of melted antimony on a sheet of paper.



c. Antimony Chlorid. Into a jar of chlorin, drop a little powdered antimony, and note the formation of antimony chlorid.



d. Stibin (Hydrogen Antimonid). Perform an experiment exactly like Experiment 48 *d*, with the exception of using antimony or antimony oxid instead of arsenic. Test the mirror with a solution of bleaching powder.

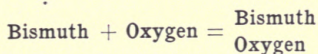


EXPERIMENT 50

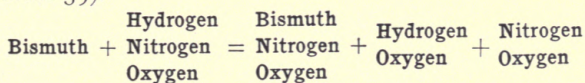
Bismuth

a. Properties. Take a piece of bismuth, and note its properties.

b. Bismuth Oxid. Oxidize a small piece of bismuth by heating it with the blowpipe on charcoal. Note the color of the oxid.

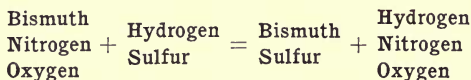


c. Bismuth Nitrate. Heat a small piece of bismuth together with a little nitric acid. The action of nitric acid upon bismuth is similar to its action upon copper (see Experiment 39).



d. Basic Bismuth Nitrate. To the solution formed in *c*, add water and note the formation of a white precipitate. This is called *basic bismuth nitrate*. (See Remark under Experiment 18.)

e. Bismuth Sulfid. Dilute a solution of bismuth nitrate; then add just enough nitric acid to prevent the formation of the basic nitrate, and pass through it a stream of hydrogen sulfid made as in Experiment 32 *c*. Note the color of the precipitate formed.

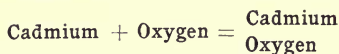


EXPERIMENT 51

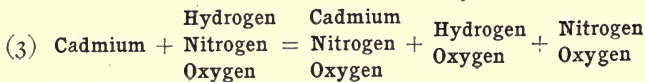
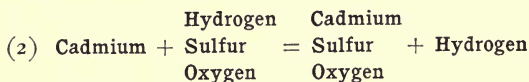
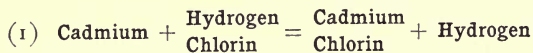
Cadmium

a. Properties. Examine a piece of cadmium, and note its properties.

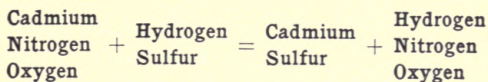
b. Cadmium Oxid. Oxidize cadmium on the charcoal, and note the color of the oxid.



c. Action of Acids upon Cadmium. Try the effect of hydrochloric, sulfuric, and nitric acids on cadmium.



d. Cadmium Sulfid. Try the effect of hydrogen sulfid upon cadmium nitrate.

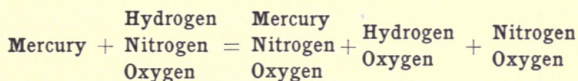


EXPERIMENT 52

Mercury

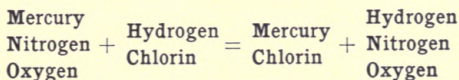
For properties and oxid see Experiment 5.

a. Mercurous and Mercuric Nitrates. Try the effect of nitric acid upon a globule of mercury.



Remark. Mercury unites with the nitrogen and oxygen of nitric acid in two proportions, the one containing the smaller quantity of the acid radical being called mercurous nitrate, and the other mercuric nitrate.

b. Mercurous Chlorid. To a solution of mercurous nitrate, add hydrochloric acid until no further precipitate is formed. The precipitate is mercurous chlorid.



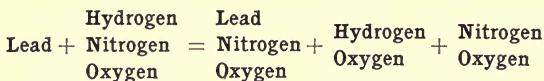
EXPERIMENT 53

Lead

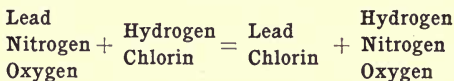
a. Properties. State the properties of lead.

b. Lead Oxid. Oxidize a small piece of lead on charcoal, and state the color of the oxid. Examine other oxids of lead.

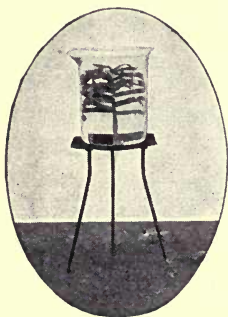
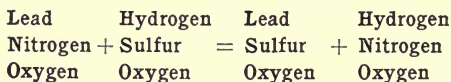
c. Action of Acids upon Lead. Try the effect of acids, hot and cold, upon lead.



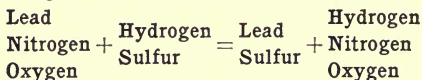
d. Lead Chlorid. To a solution of lead nitrate, add hydrochloric acid until no further precipitate is formed. Try the solubility of the resulting lead chlorid in cold and hot water.



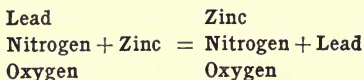
e. Lead Sulfate. To a cold solution of lead nitrate, add dilute sulfuric acid until no further lead sulfate is formed. Try its solubility in cold and hot water.



f. Lead Sulfid. Pass a stream of hydrogen sulfid through a solution of lead nitrate. State the color of the precipitate.



g. Replacement of Lead by Zinc. In a solution of lead nitrate, place a strip of zinc, and note the formation of metallic lead crystals. The zinc evidently takes the place of the lead in the nitrate. A very pretty effect can be obtained by cutting the zinc so that it can be bent into a tree-like shape.

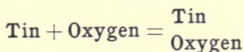


EXPERIMENT 54

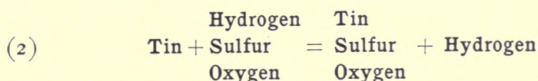
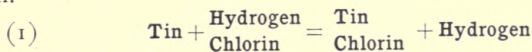
Tin

a. Properties. State the properties of tin.

b. Tin Oxid. Oxidize a little tin in a crucible by heating, stirring it with an iron rod.

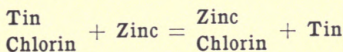


c. Action of Hydrochloric and Sulfuric Acids upon Tin. Try the action of hydrochloric* acid and sulfuric acid on tin.



d. Action of Nitric Acid upon Tin. Try nitric acid and tin. In this case, the nitrate is not formed. Tin is peculiar in that under certain conditions it can form an acid. The white precipitate you obtain is called *metastannic acid*.

e. Replacement of Tin by Zinc. In a solution of tin chlorid, place a strip of zinc, and note the metallic tin deposited on the zinc. Evidently the zinc joins the chlorin left by the tin.



EXPERIMENT 55

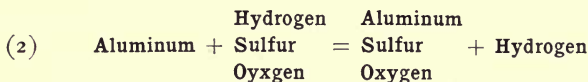
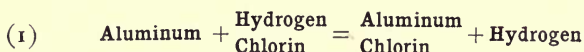
Aluminum

a. Properties. State the properties of aluminum.

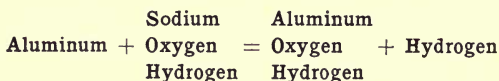
b. Does it oxidize readily?

* A small piece of copper with the tin aids the action.

c. Action of Acids upon Aluminum. Try acids, hydrochloric, nitric, and sulfuric, and verify the following : —



d. Action of Aluminum upon Alkalis. In a strong solution of sodium or potassium hydroxid, place a strip of aluminum, and note the evolution of hydrogen.

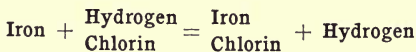


EXPERIMENT 56

Iron

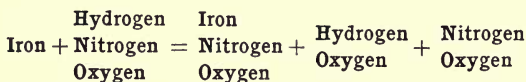
For the properties and the oxid of iron, and for the action of sulfuric acid upon iron, see Experiments 14 *a*, *b*, and 15 *f*.

a. Action of Hydrochloric Acid upon Iron. Try the effect of hydrochloric acid upon iron.



Remark. Iron unites with two proportions of chlorin, forming salts called ferrous and ferric chlorid respectively.

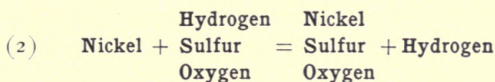
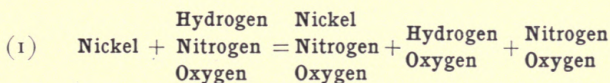
b. Try iron and nitric acid.



EXPERIMENT 57

Nickel

- a. Properties.* State the properties of nickel.
- b.* Try to oxidize nickel.
- c. Action of Acids upon Nickel.* Dissolve nickel in dilute nitric and sulfuric acids. .



EXPERIMENT 58

Barium

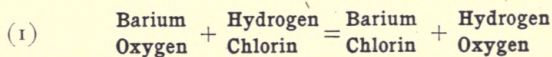
There is an element very difficult to obtain called barium. It is similar to calcium in its chemical properties, decomposing water with the evolution of hydrogen. Its compounds, however, are important.

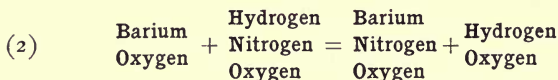
a. Barium Oxid. Examine a little barium oxid, and note its properties.

b. Barium Hydroxid. Make the hydroxid from the oxid, and test it with its moist litmus paper.

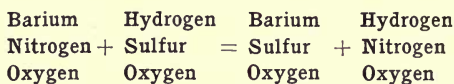
c. Barium Chlorid and Nitrate. Dissolve a little barium oxid in hydrochloric acid, and crystallize.

Dissolve a little barium oxid in nitric acid, and crystallize.





d. Barium Sulfate. To a solution of barium nitrate, add sulfuric acid until no further precipitate is formed.



e. Flame Test. In a flame, hold a platinum wire upon which there is a little barium chlorid, and note the color. This color is characteristic of barium compounds.

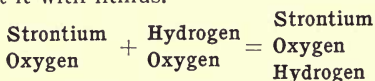
EXPERIMENT 59

Strontium

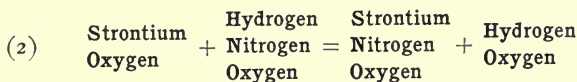
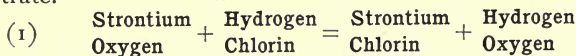
The element strontium is a yellow metal very similar to calcium and barium in its chemical properties.

a. Strontium Oxid. Get the properties of strontium oxid.

b. Strontium Hydroxid. Make the hydroxid from the oxid, and test it with litmus.



c. Strontium Chlorid and Nitrate. Make the chlorid and nitrate.



d. Flame Test. In a flame, hold a platinum wire upon which there is a little strontium chlorid, and note the color, which is characteristic of all strontium compounds.

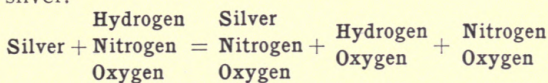
EXPERIMENT 60

Silver

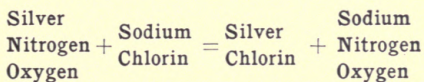
a. Properties. Examine a piece of silver, and state its properties.

b. Try to oxidize it by heat.

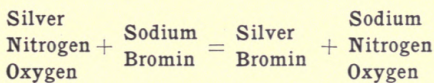
c. Action of Acids upon Silver. Try the effect of acids upon silver.



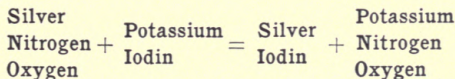
d. Silver Chlorid. Dissolve 5 grms. of silver nitrate in a little water, and add sodium chlorid until no further precipitate is formed. Allow a part of the precipitated silver chlorid to stand in the light, and try the solubility of the rest in ammonium hydroxid.



e. Silver Bromid. Perform a similar experiment, using sodium bromid instead of sodium chlorid.



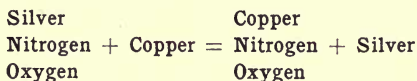
f. Silver Iodid. Perform a similar experiment, using potassium iodid.



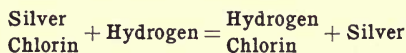
g. Silver Oxid. Perform a similar experiment, using sodium hydroxid. In this case, silver oxid is formed, instead of the hydroxid that you would naturally expect.



h. Replacement of Silver by Copper. In a solution of silver nitrate, place a strip of clean copper. Note the metallic silver formed.



i. Reduction of Silver Chlorid by means of Hydrogen. In a hard glass tube, place a little dry silver chlorid, and, passing a stream of hydrogen over it, heat it to redness. Note the metallic silver formed.



EXPERIMENT 61

Gold

a. Properties. Note the properties of gold. Hold a piece of gold leaf up to the light.

b. Action of Acids upon Gold. Try the effect of ordinary acids upon gold.

c. Action of Aqua Regia upon Gold. In a test tube, place 3 cc. of hydrochloric acid, add 1 cc. of nitric acid, and warm the mixture. Note the chlorin evolved. Place a few pieces of gold leaf on a watch glass, and add a few drops of the mixed acids (called aqua regia); then, when all is dissolved, evaporate it, and obtain crystals of gold chlorid.

The chlorin, when in the nascent state (see Experiment 40), will unite with gold, whereas, when it is already united with hydrogen in hydrochloric acid, it has no effect.

EXPERIMENT 62

Platinum

a. Properties. Note the properties of platinum. Try to melt it before the blowpipe.

b. Action of Acids upon Platinum. Try the effect of acids as you did in the case of gold.

c. Action of Metals upon Platinum. On a piece of charcoal, place a bit of lead together with a small piece of platinum, and heat them before the blowpipe. Should metals be heated in platinum crucibles?

PART II.

LAWS AND THEORIES OF CHEMISTRY

PART II.

LAWS AND THEORIES OF CHEMISTRY

Introduction. The human mind has always been more or less curious in regard to the phenomena of nature. In the time of the Greeks, the philosophers tried to explain these phenomena by speculation almost entirely. Not until the time of Bacon did men become aware of the fact that, in order to understand the changes that are going on about us in the physical world, it was necessary to make observations of these changes, and then arrange them in classes in order to see what relation they bore to each other. In this way science was born.

But it was soon found that a mere observation of facts was not enough. The question immediately arose in the mind of each observer, "How came it so?" In order to explain the facts observed, the necessity of some sort of an hypothesis was seen. Thus, when the phenomena of light and heat had been investigated by innumerable experiments, the questions arose: "What is light? What is heat?" The hypothesis was made that these phenomena were caused by the vibration of an hypothetical substance that pervaded all space. Obviously, this is a mere guess; but, so long as the facts agree with the guess, men are justified in believing that the probability in favor of its truth is great. Again, after the facts of chemistry had been sufficiently investigated by experiment, men asked themselves the question: "Why does matter act so?" The Atomic Hypothesis or Theory

was proposed to explain the facts. This theory is so closely in accord with what has been observed by experiment, and has successfully withstood so many able assailants, that men are agreed that its chance of being the truth is extremely great. However, should facts heretofore unknown be discovered that conflict with the atomic theory, we should be forced to discard it and look for some other explanation.

The student must be careful in his thinking to distinguish clearly between those things which we know from observation and experiment, and those things which we assume in order to explain what we have observed. The statements of facts learned by observation, we call *laws*. The assumptions that we make to explain these facts, we call *theories*.

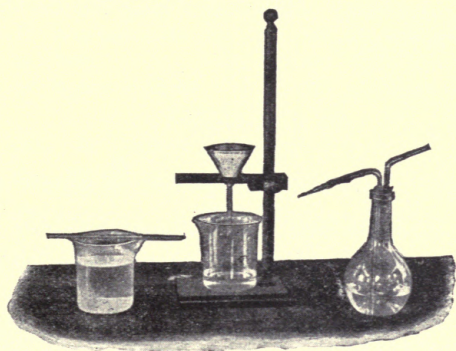
Indestructibility of Matter. The student has probably by this time become convinced, from the experiments in Part I., that matter is indestructible. This undeniable fact may be stated thus: *Matter can neither be destroyed nor created by any known human agency.* If this be true, then it must also be true that the sum of the weights of the factors of a chemical change equals the sum of the weights of the products. *Factors* signify the substances put together, and *products* the substances obtained.

EXPERIMENT I

The Sum of the Weights of Factors in a Chemical Change equals the Sum of the Weights of the Products

Weigh out exactly 5 grms. of barium nitrate and 3.5 grms. of potassium sulfate, and dissolve each portion in 100 cc. of water in separate beakers. Heat the solutions to boiling, then pour the potassium sulfate into the barium nitrate. Weigh a filter paper carefully to 1 c. gm., and, when the precipitate has settled, pour off the liquid on to

the filter, catching the filtrate in a carefully weighed porcelain dish. Pour on to the precipitate, which is still left in the beaker, about 100 cc. of hot water; then, after it has settled, pour off the liquid on to the filter as before. Do this twice, and then, by means of a fine stream of water



from a wash bottle, transfer *all* of the precipitate to the filter. A camel's hair brush will aid greatly in getting the precipitate from the sides of the beaker. If a clean glass stirring-rod is used to guide the stream, the liquid will not run down the outside of the beaker. Then wash the precipitate on the filter several times with hot water from the wash bottle. Dry the filter in an oven at a temperature of about 100° , allow it to cool, and weigh it. Evaporate the filtrate to *dryness*, and weigh. Arrange your calculations thus:—

Wt. of filter paper + barium sulfate =	
Wt. of filter paper =	
Wt. of barium sulfate =	
Wt. of dish + potassium nitrate =	
Wt. of dish =	
Wt. of potassium nitrate =	

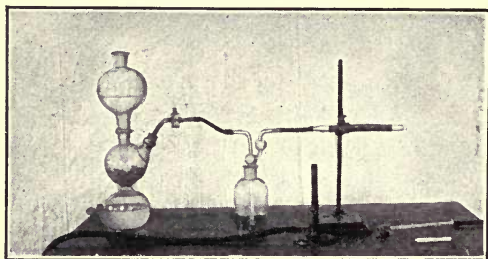
See whether the experiment shows the truth of the law.

Law of Definite Proportions by Weight. The next question that naturally arises is, "Does chemical action take place between definite quantities of elements, or are these quantities variable?" It has been found by innumerable experiments that:—*Every chemical compound has a fixed and unalterable composition.* This is known as the *Law of Definite Proportions by Weight.*

EXPERIMENT 2

Law of Definite Proportions by Weight

Weigh a porcelain boat accurately on a delicate balance. Place in it about 1 grm. of c.p. cupric oxid and weigh it



again. Then place the boat with contents in a glass tube containing stoppers, with entrance and exit tubes. Bend a piece of copper foil so that it will fit halfway around the tube for its entire length. Connect the tube with a hydrogen generator, and allow a stream of hydrogen, dried by being passed through sulfuric acid, to pass through the apparatus. When all the air has been driven out, heat the tube with a Bunsen flame until all the cupric oxid is reduced to copper. Allow the tube and contents to cool in the stream of hydro-

gen, and weigh it again. In another weighed porcelain boat, place about 2 grms. of the same oxid, and weigh again carefully. Go through the same operation, and obtain the weight of the boat and copper when all is cool.

Arrange your calculations as follows :

$$\begin{array}{rcl}
 \text{(a) Wt. of boat and oxid} & = & \text{(b) Wt. of boat and oxid} = \\
 \text{Wt. of boat} & = & \text{Wt. of boat} = \\
 \text{Wt. of copper obtained} & = & \text{Wt. of copper obtained} =
 \end{array}$$

Then make the following proportion : —

$$\begin{array}{lcl}
 \text{Wt. of oxid in a} & : & \text{Wt. of oxid in b} :: \text{Wt. of copper} \\
 \text{obtained in a} & : & \text{Wt. of copper obtained in b.}
 \end{array}$$

Law of Multiple Proportions by Weight. It will be seen that there are some elements that unite to form more than one compound, as for instance sulfur and oxygen, carbon and oxygen, and carbon and hydrogen. Analysis shows that these compounds verify the law of definite proportions. The question arises : “Is there any fixed relation between the comparative amounts of the same element in such compounds?” By experiment, chemists have proved that :

When two elements combine to form more than one compound, the amounts of one of the elements which combine with a fixed amount of the other bear to each other a simple ratio.

This is called the *Law of Multiple Proportions*.

EXPERIMENT 3

Law of Multiple Proportions

There are four compounds of copper and oxygen, and, by experiments similar to the preceding, chemists have found that the compounds have the following composition : —

	<i>Copper</i>	<i>Oxygen</i>
1.	94.09%	5.91%
2.	88.83%	11.17%
3.	79.90%	20.10%
4.	66.52%	33.47%

Find in each of the last three cases what amount of oxygen, according to the law of definite proportions, would unite with 94.09 parts of copper. Calculate thus in each case: —

$$94.09 : 88.83 :: x : 11.17.$$

Find out whether the four amounts of copper in the different cases bear to each other a simple ratio.

Quantitative Analysis. In Experiment 2, the student has performed two quantitative analyses. It is by various methods that the chemist is able to determine the percentage composition of compounds. It will perhaps be well for the student to analyze a compound by precipitating one of the elements in combination with some other element, the percentage composition of the precipitated compound being first determined.

EXPERIMENT 4

Quantitative Analysis of Sodium Chlorid

a. In order to analyze sodium chlorid, it is necessary to know the percentage composition of silver chlorid. To find this, proceed as follows: —

Weigh out on a delicate balance about .5 gm. of pure silver, recording the exact weight. Dissolve the silver in a little c.p. nitric acid diluted with water in a clean porcelain dish, and evaporate it to dryness. Dissolve the crystals in about 200 cc. of distilled water. You now have a solution of silver nitrate, which contains all the silver. Make a solution of about .5 grms. of c.p. sodium chlorid in 100 cc. of dis-

tilled water. Add from this to the silver nitrate solution as long as any precipitate is formed, and boil. When the precipitate has settled, add a drop or two of the sodium chlorid, to make sure that all the silver is precipitated. Filter off the clear liquid while it is hot. Add about 300 cc. of hot distilled water to the remaining precipitate, stir, allow it to settle, and filter it again. Do this twice, then throw all the precipitate on the filter, being sure to lose none. Wash it on the filter two or three times with hot water from the wash bottle, and allow it to dry in an oven at a temperature of about 100° . When the precipitate is dry, remove as much of it as possible to a weighed porcelain crucible. Burn the paper, holding it with a platinum wire over a clean porcelain plate. Do not let the wire come in contact with any of the precipitate. When all the carbon from the paper is burned away, transfer all the residue by means of a brush to the crucible cover. Some of the silver chlorid on the filter will have been reduced to silver by the burning paper. To change this back to silver chlorid, add two or three drops of nitric acid, warm, then add two or three drops of hydrochloric acid, and evaporate it to dryness. Now barely melt the chlorid in the crucible, allow it to cool in a desiccator, and weigh the whole.

Arrange your calculations as follows : —

Wt. of silver chlorid and crucible	=	m	=
Wt. of silver taken	=	a	=
Wt. of crucible	=	n	=
Wt. of silver chlorid	=	m - n	=

$$\text{Percentage of silver in silver chlorid} = \frac{a}{m - n} = .$$

Then $100\% - \frac{a}{m - n} \% = \text{percentage of chlorin in silver chlorid.}$

Your final result should be somewhere near 24.7%.

b. Now heat about 10 grms. of c.p. sodium chlorid in a porcelain crucible, to drive off any absorbed moisture. Cool it in a desiccator. Then weigh out on the delicate balance about .5 gm. of this, and dissolve it in about 100 cc. of distilled water in a clean beaker. Heat, and add from a silver nitrate solution until all the chlorin is precipitated as silver chlorid. Proceed exactly as in the preliminary part of this experiment, and obtain the weight of the silver chlorid.

Arrange your calculations thus:—

Wt. of sodium chlorid taken	= a	=
Wt. of crucible and silver chlorid	= b	=
Wt. of crucible	= c	=
Wt. of silver chlorid	= b - c	=

Since silver chlorid contains 24.73% of chlorin, there are 24.73 ($b - c$) grms. of chlorin in the silver chlorid obtained. But this chlorin came from the a grms. of sodium chlorid taken; therefore the percentage of chlorin in sodium chlorid is $\frac{24.73 (b - c)}{a}$. Your result should be about 60.6%.

Atomic Theory. As soon as the preceding facts, i.e., the indestructibility of matter, the law of definite proportions, and the law of multiple proportions, had been established, men of science immediately began to look for some explanation of them. It remained for Dalton in the early part of this century to propose an hypothesis that is accepted to the present day. We are compelled to believe in its truth, for we can think of no other hypothesis that will account for the facts.

Dalton's hypothesis is this. All matter is composed of minute particles, which cannot be divided by any chemical means. Each of these particles of any simple substance (i.e. element) is like every other one of that sub-

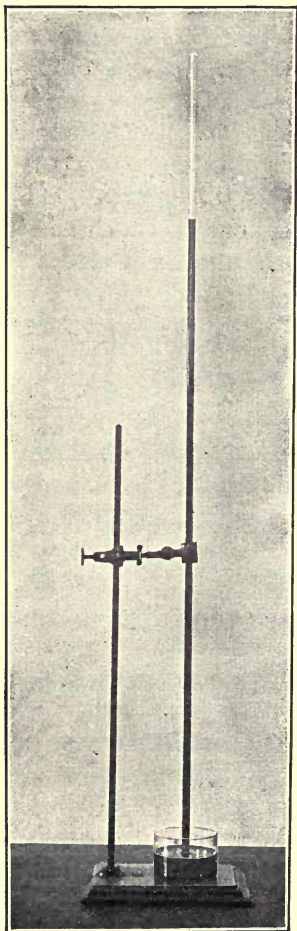
stance, both in *properties* and *weight*. These particles are called atoms. Therefore, *an atom is the smallest particle of matter that can enter into chemical combination*. These like or unlike atoms may unite with each other, and form other particles which are called molecules. Therefore, *a molecule is the smallest particle of matter that can exist and still have all the properties of the substance, and is made up of atoms chemically united*.* This, briefly stated, is Dalton's Atomic Theory. It must be borne in mind that it is simply an hypothesis or guess formulated to explain known facts. If the hypothesis *is* the actual truth, then it follows that, in the case of copper and oxygen, when their atoms are chemically united to form copper oxid, the proportions will always be the same. In the case of two compounds of the same elements, as for instance the two oxids of copper, it follows that the amount by weight of oxygen in the first must bear a simple ratio to the amount by weight of oxygen in the second. For evidently twice as many atoms of oxygen would unite with a definite number of atoms of copper to form the second oxid as would unite with the same number of atoms of copper to form the first oxid. In all cases, the ratio must be a simple one, since, if matter is made up of atoms, chemical combination must take place between the whole atoms.

Atmospheric Pressure. Barometer. In order to understand the full meaning of the atomic theory, it will be necessary for the student to perform a number of experiments with gases. To do this, he must understand the meaning of atmospheric pressure, and the effects of changes of pressure and of temperature on the volumes of gases. First let us take up the subject of atmospheric pressure.

* Some molecules are made up of single atoms.

EXPERIMENT 5

Barometer



Take a glass tube of uniform bore (5 mm.) about a meter long, and close one end. Fill it with mercury, and, by means of a wire, remove what air bubbles cling to the sides of the tube. Invert the tube in a trough of mercury, and measure the height of the column. Such an instrument as this is called a *barometer*. Now since the mercury remains in the tube, some force must be exerted to hold it there, and the only thing that can do so is the air pressing down upon the surface of the mercury in the trough. Therefore the weight of the mercury in the tube balances a column of air of equal cross section extending to the top of the atmosphere. If you should perform this experiment on different days, you would observe that the column stands at various heights. Hence we see that the pressure or weight of the atmosphere resting upon a certain part of the earth's surface varies at

different times. The standard pressure is the pressure that will hold up a column of mercury 760 mm. high, when the thermometer stands at 0° C. This is about 15 lbs. to the square inch or 1033.6 grms. to the square centimeter.

Boyle's Law. Now we are ready to investigate the effect of pressure upon the volume of a gas. We all know by every day experience that gases can be reduced to a smaller volume by increasing the pressure upon them, and that they expand when the pressure is removed; but most of us are perhaps ignorant, unless we have studied physics, of the exact effect of the pressure upon the volume. The following experiment will show that *the volume of a gas varies inversely as the pressure exerted upon it*; i. e., if you double the pressure, you divide the volume by two, or if you treble the pressure, you divide the volume by three.

Expressing the law algebraically, we have

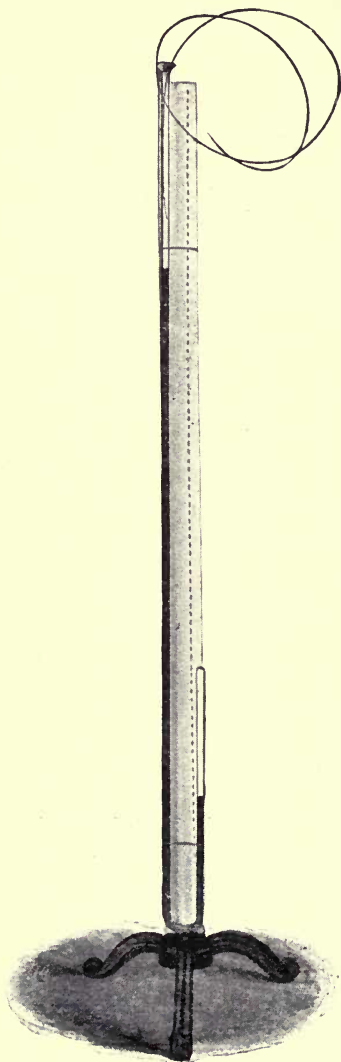
$$V : V' :: P' : P$$

in which V is the volume at the pressure P , and V' is the volume at the pressure P' . This is called *Boyle's Law*.

EXPERIMENT 6

Law of Boyle

Have ready a clean dry glass tube of about 6 mm. bore, closed at one end and bent so as to form a narrow letter J, the hook being about 30 cm. long. Let the long arm be about 100 cm. in length. Into this, pour a little clean mercury, so that it stands about the same height in both arms. It will do no harm if it stands a little higher in the long arm. Now fasten the tube to a perpendicular support, and measure from the base of the support to the top of the bore in



the short arm, reading the scale to tenths of millimeters. The edge of a visiting card held across the tube will aid the eye in taking the reading. Then measure the height of the mercury in the long and short arms. Do not touch the column of confined air with the hand; for, in that case, the warmth of the hand will cause the air to expand and thus give incorrect readings on the scale. Now take the reading of the barometer. Pour into the tube enough mercury to make the column in the long arm stand about 10 cm. higher. To remove any air bubbles that may be held imprisoned by the mercury, push a long iron wire down into the mercury; then withdraw it, at the same time turning it around in the tube. Take the measurements as before. Do this as many times as the length of the tube permits.

Arrange your numbers in a table thus : —

Distance from base to top of bore, short arm = A.	Height of mer- cury in short arm = B.	Height of mer- cury in long arm = C.	Barometer read- ing = H.	Height of mer- cury in long arm exerting pressure = C - B.	Total pressure = H + C - B.	Length of air column = A - B.

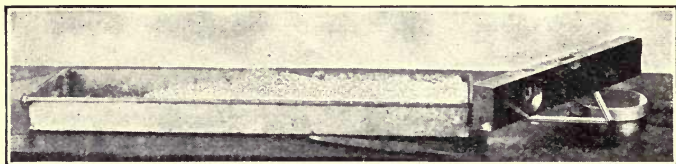
Compare the volume at any pressure with the volume at any other pressure, and verify the proportion $V : V' :: P' : P$.

EXPERIMENT 7

Law of Charles

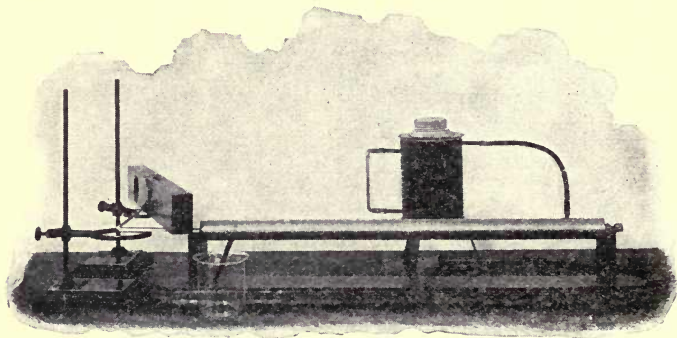
Have ready a tube of 1 mm. bore about 100 cm. long, bent at right angles at a distance of about 50 cm. from the closed end, and containing a column of dry air confined by means of mercury, the mercury extending somewhat beyond the bend. Insert this through a cork stopper in a hole in the end of a shallow tin tray containing melting ice. Allow the air column to remain in the ice for a few minutes, keeping the ends of the mercury column at the same level.

Mark the end of the air column, and measure with a meter rod. Now insert the confined column of air in a steam jacket containing a thermometer; then allow steam to pass through the apparatus. Keeping the two ends of the mercury level, again mark the end of the air column. Remove it



from the jacket, and measure it. The difference between the lengths in ice and in steam will be the amount by which the air column has expanded in being heated 100° .

To find the amount one centimeter expands for one degree, divide this result by the original length \times the number



of degrees heated. This should give a number near .00366, which reduced to a fraction equals about $\frac{1}{273}$. We therefore see that, *for every degree it is heated at a constant pressure, a volume of air expands $\frac{1}{273}$ of what its volume was at 0° C.* This is known as the *Law of Charles*.

Theoretically then, the volume of a gas at 273° below zero would become nothing. Of course this is impossible, but the point -273 is taken as the absolute zero of temperature. Then a° in the ordinary scale would be $a + 273^{\circ}$ in the absolute scale.

Deduction of the Formula.

Let A be the volume of a gas at 0° ,

then $A + \frac{A}{273}$ will be the volume at 1°

$A + \frac{2A}{273}$ will be the volume at 2°

$A + \frac{3A}{273}$ will be the volume at 3°

$A + \frac{tA}{273}$ will be the volume at t°

and $A + \frac{t'A}{273}$ will be the volume at t'°

Representing the volume at t° by V , and the volume at t'° by V' , we have

$$V = A + \frac{tA}{273}, \text{ and } V' = A + \frac{t'A}{273}$$

Dividing and canceling the A 's, we have

$$\frac{V}{V'} = \frac{273 + t}{273 + t'}$$

Combining this with the formula for pressure, $V : V' :: P' : P$, we obtain

$$\frac{VP}{273 + t} = \frac{V'P'}{273 + t'}$$

Boyle's and Charles's Laws are General. In the experiments we have used air, but it will be found by trial that other gases act in the same way as air. Thus we see that different kinds of matter in the gaseous state have a property in common.

Examples. 1. The volume of a gas at 765 mm. pressure and 20° C. is 450 cc.; what will its volume be at 0° and 760 mm.? Ans. 422. +

2. At what temperature, pressure remaining the same, will a gas be double its volume at 0° ? Ans. 273° .

3. Reduce the following volumes with the annexed temperatures and pressures to volumes at second temperatures and pressures:—

V	t	P	V'	t'	P'
500	100	800 mm.	?	20.	650 mm.
1000	70	900 mm.	?	0.	200 mm.

EXPERIMENT 8

Weight of a Liter of Air

Fit a strong glass bottle of about one-liter capacity, with a one-holed rubber stopper containing a glass tube. Wire on to the tube a short piece of antimony-rubber tubing fitted with a pinch cock. Make all connections air-tight with vaseline. Weigh the bottle and connections *carefully* to one centigram. Attach an air pump to the rubber tube, pump out as much air as possible, close the tube with the pinch cock, and weigh the bottle carefully again. Open the tube under water of about the same temperature as the room. After the water has ceased rushing in, hold the bottle so that the water stands at the same level both inside and outside, close the pinch cock, remove the bottle, and weigh it again. Take the readings of thermometer and barometer.

Arrange your calculations thus:

Wt. of bottle full of air	= a
Wt. of bottle, air pumped out	= b
Wt. of air pumped out	= a - b
Wt. of bottle with water	= d
Wt. of water = vol. air pumped out	= d - a = c.

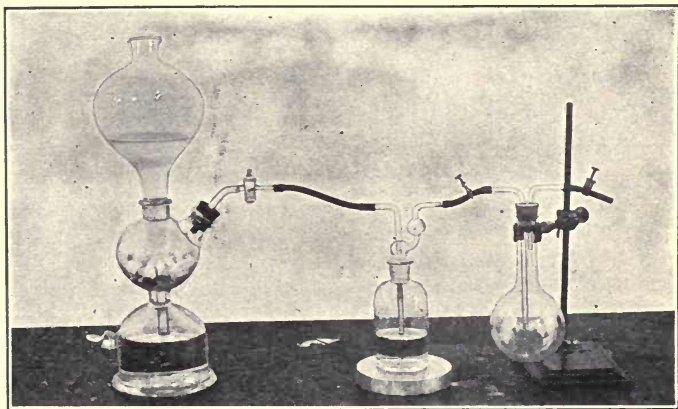
Therefore c cc. of air at the temperature and pressure on the day of the experiment weigh $a - b$ grams. Calculate the weight of one liter. You should obtain with this apparatus somewhere near 1.2 or 1.3 grams.

The correct weight of a liter of air at 0° and 760 mm. pressure is 1.293 grams, which number we shall hereafter use.

EXPERIMENT 9

Weight of a Liter of Carbon Dioxid at 0° and 760 mm.

Fit a dry one-liter flask with a two-holed rubber stopper. In the holes, insert glass tubes bent at right angles, one reaching to the bottom of the flask and the other just



through the stopper. Fit with rubber tubes and pinch cocks. Make all joints air-tight with vaseline, and weigh carefully to one centigram. Allow carbon dioxid, dried by being passed through sulfuric acid, to pass through the flask until it is completely filled. Close the tubes and disconnect them. Open

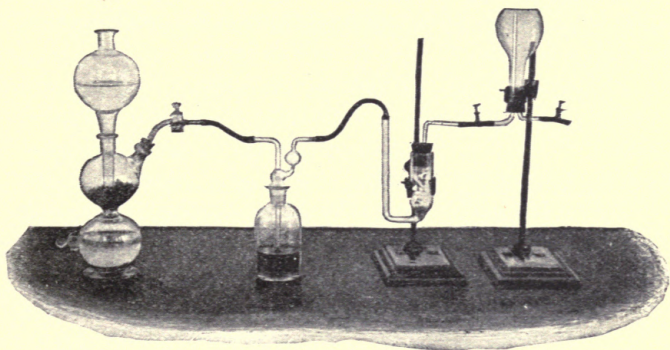
one pinch cock for a moment to relieve the pressure, close it again, and weigh the flask. Take the temperature near the flask and the reading of the barometer. The increase in weight will be the amount by which the weight of the flask full of carbon dioxid exceeds the weight of the flask full of air. Mark the capacity of the flask, fill it with water to the mark, and weigh. Since 1 cc. of water weighs one gram, the number of grams of water in the flask equals the number of cubic centimeters in the volume of the flask. Calculate what this volume of air at the temperature and pressure at the time of the experiment would be at 0° and 760 mm. Knowing that one liter of air at 0° and 760 mm. weighs 1.293 grams, find the weight of the air. Add to this the increase in weight due to the carbon dioxid, and obtain the weight of the same number of cc. of the gas. From this result, obtain the weight of 1000 cc., i. e., 1 liter of carbon dioxid. You should obtain a number somewhere near 1.97.

EXPERIMENT 10

Weight of a Liter of Hydrogen

Perform a similar experiment with hydrogen gas. Use a 100 cc. flask, and weigh it on a delicate balance. Keep the flask inverted when filling, and be sure the hydrogen is *dry*. To insure this, pass the hydrogen not only through sulfuric acid but also through a U tube containing pieces of granular calcium chlorid. Make the hydrogen with zinc and sulfuric acid (one part acid to five of water). In this case, the weight of the flask full of hydrogen is less than that of the flask full of air; hence you must *subtract* the decrease from the calculated weight of the flask full of air at 0° and 760 mm. You should obtain a result somewhere near .09 grams for 1 liter. (See illustration on page 107.)

Density. By the density of an element or compound is meant its weight in gaseous form compared with the weight of an equal volume of hydrogen, the temperature and pressure being the same.



EXPERIMENT 11

Calculate the density of air and carbon dioxide compared with hydrogen, and show that they are 14.43 and 22 respectively.

EXPERIMENT 12

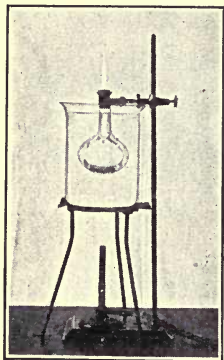
Density of a Liquid in the form of Vapor

Have ready a clean dry 100 cc. flask fitted with a one-holed rubber stopper containing a glass tube drawn out to a diameter of about 1 mm. Make it air-tight with a little vaseline, and weigh it carefully on a delicate balance. Take the reading of the thermometer and barometer. Place about 20 cc. of alcohol in the flask, and put the flask into boiling water up to its neck. When the alcohol has boiled away,*

* Do this gradually so as not to blow out the stopper.

ignite the vapor issuing from the tube, and the moment that the flame dies away, seal the tube with a blowpipe. Now weigh the flask again. Fill the flask with water by breaking the end of the tube under water. If the flask does not *fill* almost completely, the experiment must be repeated. The weight of the flask full of water minus the weight of the flask gives the volume.

The increase in weight will be the weight by which the flask full of vapor of alcohol at 100° exceeds the weight of the flask full of air at the temperature of the room. Find what the volume of the air in the flask would become at 0° and 760 mm. Knowing the weight of a centimeter of air, find the weight of this volume of air. Add the increase in weight due to the alcohol vapor, and obtain the weight of the alcohol vapor at 100° and the pressure at the time of the experiment. Calculate the volume of the air in the flask at 100° . Let us call this M . Then if V is the volume



of the flask at the temperature of the experiment, the weight of the air left in the flask at 100° would be $\frac{V}{M}$ times its weight at the temperature of the experiment. We now have the weights of the flask full of air and alcohol at 100° . Hence the density of alcohol referred to air is easily found. But air is 14.43 times as heavy as hydrogen; therefore, to find the density of alcohol compared with hydrogen, multiply by 14.43.

Weights of the Atoms. The student will remember that, in the atomic theory, it is assumed that the atoms of each ele-

ment are of the same weight. Under the present condition of methods of investigation, it is impossible to determine the absolute weight of these atoms with any satisfactory degree of accuracy. However, as soon as the atomic theory was accepted, men began to try to find out the weight of these atoms in terms of some atom taken as a standard. Hydrogen being the lightest element known, its atom was taken as this standard. It required a number of years and the establishment of auxiliary principles to fix even these numbers with any degree of certainty. As time went on, investigation brought out the principles of *chemical, electrical, thermal, and isomorphic*, equivalence of elements. The *Law of Definite Properties by Volume*, *Avogadro's Law*, and the *Periodicity of the Elements*, were also established; and with the aid of these it has been possible to determine the relative weights of the atoms to a remarkable degree of accuracy.

We will take up these subjects in turn.

Chemical Equivalence. By experiment, it is found that on causing hydrogen and chlorine to unite, the ratio of the parts by weight are 1:35.4. That is, the chemical value of chlorine compared with hydrogen is 35.4. In like manner, 8 parts by weight of oxygen unite with 1 part by weight of hydrogen. We therefore say the *chemical equivalence* of oxygen is 8. This is also called its *combining number*. To find the chemical equivalent of an element in terms of another, it is necessary only to find the amount by weight of the first that unites with a fixed amount of the second, or the amount which replaces a fixed quantity in a compound. In some cases, that is, where there is more than one compound between the same element, it is difficult to decide which number to take. If we knew the exact number of atoms of each kind that unite to form a compound, or the exact number of atoms of an element replaced in a compound

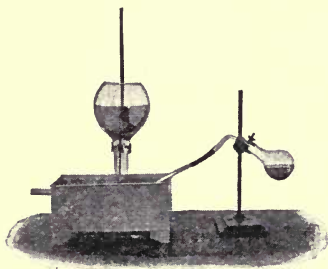
by one atom of another element, it would be easy to determine the atomic weight directly from the chemical equivalence. Thus, if we knew that one atom of hydrogen united with one atom of oxygen to form water, we could be sure that the atomic weight of oxygen was 8. Or, if we knew that two atoms of hydrogen united with one of oxygen, we could be sure of the number 16 as the atomic weight of oxygen. Not knowing this, however, we see that our data is insufficient, and we must look farther.

EXPERIMENT 13

Chemical Equivalence of Zinc

Have ready a clean 100 cc. flask containing a dilute solution of c.p. hydrochloric acid (10 cc. of water to 20 cc. of acid). Fit the flask with an air-tight one-holed rubber stopper fitted with a delivery tube. Weigh out as nearly as possible 1.25

grms. of c.p. zinc, recording the exact weight. Place in a pneumatic trough a 500 cc. flask inverted and full of water. Have the delivery tube clamped so that its end is under the mouth of the flask. Remove the stopper from the 100 cc. flask, drop the zinc into it, close it



quickly, and catch *all* the evolved hydrogen. When all the zinc is dissolved, hold the flask containing the hydrogen so that the water stands at the same level both inside and outside. Do not touch with the hand that part of the flask containing the hydrogen. Take the temperature near the flask

and the reading of the barometer. Mark the point on the flask at which the gas stands, remove it, and fill it with water to this point. Then weigh the flask and water. From this, obtain the volume of the gas.

To obtain the best results, it will be necessary to take into account the fact that, besides the hydrogen, there is vapor of water in the flask, which helps bear the pressure of the air upon the surface of the water in the trough. This varies for different temperatures. The following table is near enough for all practical purposes.

DEGREES.	MILLIMETERS OF MERCURY.	DEGREES.	MILLIMETERS OF MERCURY.	DEGREES.	MILLIMETERS OF MERCURY.	DEGREES.	MILLIMETERS OF MERCURY.
10	9.2	16	13.6	22	19.7	28	28.1
11	9.8	17	14.4	23	20.9	29	29.8
12	10.5	18	15.4	24	22.2	30	31.5
13	11.1	19	16.4	25	23.5	31	33.4
14	11.9	20	17.4	26	25.0	32	35.4
15	12.7	21	18.4	27	26.5	33	37.4

Subtract the number of millimeters of mercury in the table at the temperature corresponding to the temperature observed near the flask, and you will have the true pressure to which the hydrogen is subjected.

Calculate by means of the formula

$$\frac{V P}{273 + t} = \frac{V' P'}{273 + t'}$$

what the volume of the hydrogen would be at 0° and 760 mm. Knowing that 1 cc. of hydrogen, at 0° and 760 mm., weighs .00009 grms., calculate the weight of the hydrogen; then obtain from this the number of grams of zinc that will replace one gram of hydrogen. You should obtain somewhere near the number 32.5.

Should time allow, it will be interesting for the student to determine the number for iron, using piano wire instead of zinc. The number in this case is 27.9.

Electrical Equivalents. Michael Faraday found that, when a current of electricity was passed through a substance which it decomposed, the quantity of the substance decomposed varied with the strength of the current. If the same current passes through two substances, such as a solution of copper sulfate and a solution of silver nitrate, the amount of copper deposited is to the amount of silver deposited as the chemical equivalent of copper is to the chemical equivalent of silver. But it is found that the same difficulty exists as with finding the chemical equivalents; namely, that, if there are two or more different compounds between the same elements, then as many different electrical equivalents are obtained. Thus copper gives 31.6 and 63.2, and mercury gives 99.9 and 199.8. So we have still the same question as to which number shall be taken as the atomic weight.

EXPERIMENT 14

Electrical Equivalents

Have ready two Daniel cells. In a beaker, put a solution of copper sulfate, and in another, put a solution of nickel

sulfate. Cut out two copper plates three centimeters wide by six long, making a hole in each, and weigh one carefully. Cut out two similar nickel plates of the same size, and weigh one. Place the copper plates in the copper sulfate, and the nickel plates in the nickel sulfate. Connect the copper plate that was not weighed with the weighed nickel plate by means of a wire, and then put the two solutions in circuit with the two Daniel cells, so that the weighed plates are the cathodes. Allow the current to pass for about twenty minutes; then disconnect, dry, and weigh the plates again.



Arrange your calculations as follows: —

Wt. cathode copper plate	=
Wt. after	=
Difference	=
Wt. cathode nickel plate	=
Wt. after	=
Difference	=

Compare the weights of the metals deposited, and see whether they are in the same ratio as the chemical equivalents of the two metals, 63.2 and 58.6.

Specific Heat. We have noticed in our every day experience that it requires a different amount of heat to raise the temperature of various substances the same number of degrees. Since heat is a form of energy, it can be measured. To measure a quantity of heat, it will be necessary to use some definite quantity as a unit of measurement. The unit used is the *calorie*. *A calorie is that quantity of heat that is used up in raising one gram of water one degree centigrade.*

Therefore to raise one gram of water 2° , or two grams 1° , requires two calories. *The amount of heat, measured in calories, necessary to raise one gram of a substance 1° C., is called the specific heat of that substance.* Below is given the specific heat of a number of the elements.

Iron .114	Copper .094	Silver .057	Gold .032
Nickel .108	Zinc .095	Tin .055	Lead .031

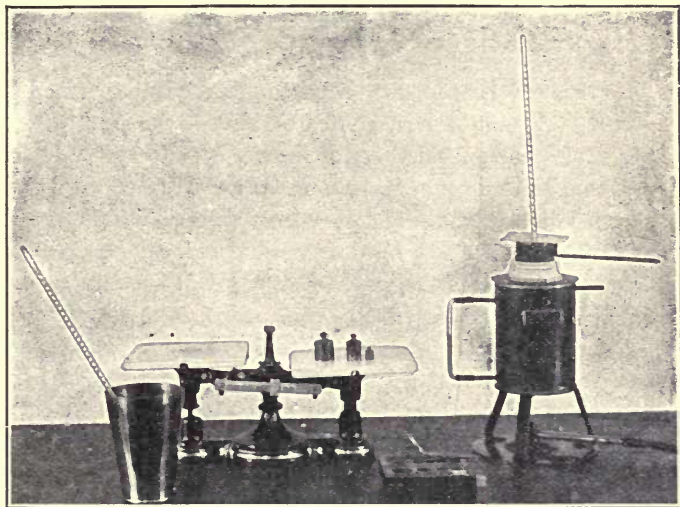
Apparatus used in Determining Specific Heat. To find the specific heat of elements experimentally, two pieces of apparatus are necessary; first, a vessel, called a calorimeter, to hold the substance. An ordinary nickel-plated lemonade shaker will answer very well. Secondly, we must have some sort of apparatus in which the substance can be heated to a constant temperature. Such an apparatus (see illustration, p. 115) can be obtained at small cost from dealers in laboratory supplies, or one can be set up as follows. Have ready a large beaker of about 1000 cc. capacity, one-quarter full of water. Then cut a thin board in the shape of a ladle, large enough to make the broad part cover the large beaker. Cut a hole in the center of the broad part large enough to hold another beaker of about 200 cc. capacity.

EXPERIMENT 15

Specific Heat of Lead

To determine the specific heat of lead with any degree of accuracy by means of the apparatus here used, we shall assume that the specific heat of the material of the calorimeter is known. In the case of brass it is .094, in the case of glass .2.

Weigh out 500 grms. of fine shot, and place the shot in the small beaker. Set the small beaker in the large beaker, and heat the water in the large beaker to boiling. Cover the small beaker with a piece of cardboard. Meanwhile weigh the calorimeter, and add exactly 100 grms. of water



cooled about 8° below the temperature of the room. Stir the shot occasionally and thoroughly with a thermometer, and note when the temperature becomes constant. Then pour the shot quickly into the calorimeter, stirring with another thermometer, and note the temperature of the mixture. (The illustration shows apparatus mentioned on p. 114.)

Let us see what has taken place. The amount of heat lost by the shot (we neglect what little heat has been radiated off) must be equal to the amount of heat gained by the water and the calorimeter.

Arrange your calculations as follows: —

Weight of the shot	=	s	=
Weight of the calorimeter	=	c	=
Weight of the water	=	w	=
Temperature of the shot	=	t_s	=
Temperature of the water	=	t_w	=
Temperature of the mixture	=	t_m	=
Specific heat of substance of cal.	=	h	= .094

Amount of heat lost by shot = amount of heat gained by water + amount of heat gained by the calorimeter.

Let x = the specific heat of lead.

Then, since 1 gram of lead loses x calories in falling 1° , s grams of shot will lose $s(t_s - t_m)x$ calories in falling $t_s^\circ - t_m^\circ$. Since one gram of water gains one calorie in being heated one degree, w grams of water will gain $w(t_m - t_w)$ calories in being raised $t_m^\circ - t_w^\circ$. Since one gram of the substance of the calorimeter gains h calories in being raised one degree, c grams will gain $ch(t_m - t_w)$ calories in being raised $t_m^\circ - t_w^\circ$. Therefore we have the original statement expressed algebraically.

$$s(t_s - t_m)x = w(t_m - t_w) + ch(t_m - t_w)$$

Substituting your observed values, you should obtain, by solving for x , a number somewhere near .031 as the specific heat of lead.

EXPERIMENT 16

Specific Heat of Iron

In the same way, find the specific heat of iron, using iron filings free from grease or oil. You should obtain somewhere near the number .114.

Thermal Equivalents. In 1819, two French chemists, Dulong and Petit, noticed that there existed a simple relation between the chemical equivalents of the elements and their specific heats; namely, that the product of the two numbers always approximates 6.4, or half that number. In most cases it is 6.4. The following table shows this remarkable relation. The chemical equivalents are doubled where it is necessary to obtain approximately the product 6.4.*

ELEMENT.	ATOMIC WEIGHT.	SPECIFIC HEAT.	PRODUCT.	ELEMENT.	ATOMIC WEIGHT.	SPECIFIC HEAT.	PRODUCT.
Iron	55.6	.114	6.4	Silver	107.11	.057	6.1
Nickel	58.24	.108	6.3	Tin	118.15	.055	6.5
Copper	63.12	.094	6.0	Gold	195.74	.032	6.3
Zinc	64.91	.095	6.1	Lead	205.36	.031	6.4

The product of the atomic weight and the specific heat varies somewhat, of course; but it is only reasonable to suppose that, if our numbers were absolutely correct, this product would always be the same. The quotient of the number in the last column divided by the specific heat is called the *thermal equivalent*.

Explanation. Here we have a method of aiding us in determining what numbers we shall take for the atomic weights. The only way we can explain the relation noted in the last paragraph is by saying that *all atoms have the*

* This product will probably be somewhat smaller than 6.4 as the atomic weights are more and more accurately determined.

same capacity for heat; that is, *it requires the same amount of heat to raise every atom, no matter of what kind, one degree in temperature.* For example, the atomic weight of lead is 205 microcriths (a microcrith being the weight of an atom of hydrogen), and that of iron is 56 microcriths. If we assume that it requires the same amount of heat to raise an atom of each one degree, it will take $\frac{1}{205}$ as much to raise one microcrith of lead one degree, and $\frac{1}{56}$ as much to raise one microcrith of iron one degree. But the atomic weights must bear to each other the same ratio as their mass weights, therefore the amount of heat necessary to raise one gram of lead one degree must be to the amount of heat necessary to raise one gram of iron one degree as $\frac{1}{205}$ is to $\frac{1}{56}$, or as 56 is to 205. The specific heats of lead and iron, .031 and .114 respectively, are to each other in this ratio. We are then justified in believing that our assumption — that every atom requires the same amount of heat to raise it one degree in temperature — is true.

We have therefore considerable evidence to warrant us in believing that

$$\text{atomic weight} \times \text{specific heat} = \text{a constant, i.e., } 6.4$$

From this we have

$$\text{atomic weight} = \frac{6.4}{\text{specific heat}}$$

or

$$\text{specific heat} = \frac{6.4}{\text{atomic weight}}$$

From Experiment 13, we find that 32.5 grams of zinc replace one gram of hydrogen. Suppose we take 32.5 for the atomic weight, and divide 6.4 by this number. We obtain .193 for the specific heat. Now let us double the number 32.5, and divide 6.4 by the result. We then obtain .096 for the specific heat. Since the specific heat of zinc is

found by experiment to be .0955, we are justified in believing that the atomic weight is 65, just *twice* its chemical equivalence.

Isomorphic or Crystallographic Equivalents. A short time after Dulong and Petit made their discovery, Mitscherlich, a German chemist, found that certain elements can replace others in a compound without changing the crystalline form. Such elements are said to be *isomorphous*. The replacement always takes place in definite quantities. These quantities are called the *crystallographic equivalents*.

The chemical equivalence of silver is 107. In a crystal, 107 parts by weight of silver can be made indirectly to replace 65 parts of zinc; hence we say that 107 parts of silver is the crystallographic equivalent of 65 parts of zinc. But the chemical equivalence of zinc is 32.5. The only explanation we can give of this is that one atom of zinc replaces two atoms of hydrogen; and hence the atomic weight of zinc is not the same as its chemical equivalence, but twice 32.5, or 65. This is evidently in harmony with what we learned in regard to the atomic weight of zinc from the Law of Dulong and Petit. In like manner, the atomic weights of other elements have been investigated, and it has been found that this principle is a great aid in determining the right number to use in many cases. It has not however been found to be universally true.

EXPERIMENT 17

Law of Definite Proportions by Volume

Refer to Experiment 6 *e*, Part I., and note the respective volumes of hydrogen and oxygen that unite to form water. From Experiment 24 *b*, Part I., note the number of volumes

of hydrogen that unite with one volume of chlorin to form two volumes of hydrochloric acid gas.

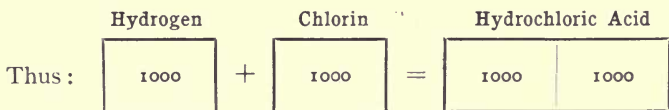
Statement of the Law. From such experiments as the preceding, it has been found that, *when gases unite to form compounds, the relative volumes bear to each other a simple ratio.* This is called the Law of Definite Proportions by Volume, or the Law of Gay Lussac.

Avogadro's Hypothesis. In order to explain the above fact, it will be necessary for us to make the hypothesis that *equal volumes of gases, temperature and pressure being the same, contain an equal number of molecules.* This is called the Hypothesis of Avogadro. (Avogadro was an Italian physicist of the early part of the nineteenth century.) That this assumption is true, is indicated by both physical and chemical considerations. For instance, we know that all gases, under changes of pressure, act in the same manner; this is also true for changes of temperature. Besides, by accepting the atomic theory and arguing from a purely mathematical standpoint, the truth of this hypothesis has been shown to be an absolute necessity.

Molecular Weights of Gaseous Elements and Compounds. It must follow then, if equal volumes of gases contain an equal number of molecules, that the ratio of the weight of a molecule of an element or compound to the weight of a molecule of hydrogen is the same as the ratio of the weights of equal volumes of the two gases; that is, the same as the density of the element or compound. Thus the molecular weight of carbon dioxid compared to the weight of a *molecule* of hydrogen must be $\frac{1.97}{.09} = 22$. It must be borne in mind, however, that we have not yet determined how many atoms there are in the hydrogen molecule. When we have done that, we can obtain the molecular

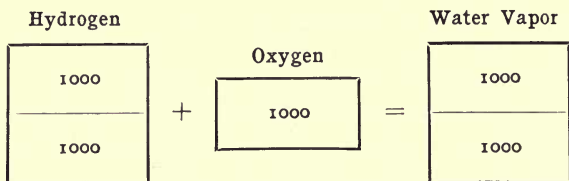
weight in terms of the hydrogen *atom*. This we proceed to do.

Number of Atoms in the Hydrogen Molecule. We shall be compelled to believe that a molecule of hydrogen contains *at least* two atoms. For we know by Experiment 24 *b*, Part I., that one volume of hydrogen unites with one volume of chlorin to form two volumes of hydrochloric acid gas. Let us suppose that in the one volume of hydrogen there are 1000 molecules, then from Avogadro's Hypothesis there must be 1000 molecules in the one volume of chlorin, and 2000 molecules in the two volumes of hydrochloric acid.



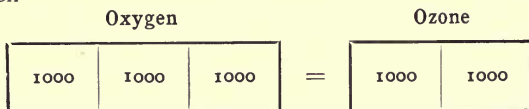
But every hydrochloric acid molecule is shown by analysis to be made up of hydrogen and chlorin; therefore in the 2000 molecules of hydrochloric acid there must be *at least* 2000 atoms of hydrogen. Now these 2000 atoms of hydrogen were originally in the 1000 molecules of hydrogen; therefore in these 1000 molecules there must be *at least* 2000 atoms, and thus each molecule must contain at least two atoms. In all the chemical investigations that have been made, nothing has ever been found that would seem to indicate that the hydrogen molecule contains *more* than two atoms; so we shall accept this number. Likewise each molecule of chlorin must contain at least two atoms.

Number of Atoms in the Oxygen Molecule. It is proved by experiment that two volumes of hydrogen unite with one volume of oxygen to form two volumes of water vapor. Suppose, as before, that one volume contains 1000 molecules. We then have



Each of the 2000 molecules of water vapor must contain at least one atom of oxygen; but these 2000 atoms came from the 1000 molecules of oxygen; hence each oxygen molecule must contain at least two atoms. The only thing that makes us believe oxygen ever contains more than two atoms is the fact that, when a silent electric discharge is passed through dry oxygen gas, it decreases to two-thirds of its former volume. This is easily explained in the light of Avogadro's Hypothesis. Suppose, as before, one volume contains 1000 molecules.

Then



Nothing is simpler than to believe that three oxygen molecules, each containing two atoms, have been broken up, and then reunited to form two molecules each containing three atoms. Hence ozone differs from oxygen in that it contains three atoms to the molecule instead of two. Thus we see that we have a very rational explanation of allotropic forms.

Atomic Weight of Oxygen. In the light of what we have seen, it follows that *the molecular weight of any gaseous compound may be obtained by multiplying its density by two.* Thus, since a molecule of carbon dioxide is 22 times as heavy as a molecule of hydrogen, and since a molecule of

hydrogen contains two atoms, therefore a molecule of carbon dioxid will be 44 (i.e., 2×22) times as heavy as an *atom* of hydrogen. In like manner, the molecular weight of oxygen is found to be 32; and, since there are two atoms in the molecule, its atomic weight is 16.

Molecular Weights of Compounds not Capable of Being Volatilized. To solid compounds not capable of being volatilized, we are unable to apply the ordinary method of finding the molecular weight. We are obliged to attack the question in another way. This is best illustrated by the following experiment.

EXPERIMENT 18

Molecular Weight of Potassium Chlorate

Dry in an oven at about 150° a quantity of powdered c.p. potassium chlorate. Weigh a porcelain crucible and its cover on the delicate balance. Then weigh out in the crucible about 2 grms. of the dry chlorate. Heat it on a triangle, gently at first, holding the cover over the crucible by means of a pair of long forceps in order to catch any of the melted substance that may spatter. Be careful not to heat too strongly at first, since the chlorate may bubble over the sides of the crucible and spoil the experiment. When all the oxygen is driven off, cool the crucible in a desiccator and weigh it again. Continue to heat and weigh it, until you obtain a constant weight. From the numbers obtained, it is possible to get the molecular weight.

We know that potassium chlorate is composed of potassium, chlorine, and oxygen. If we knew how many atoms of oxygen there are in the potassium chlorate molecule, we could make our calculation easily. Now chemists have de-

terminated (it is a problem beyond elementary students) that a molecule of potassium chlorate must contain either three or a multiple of three atoms of oxygen. Three has been chosen as the more probable number. If then there are three atoms of oxygen in the molecule, the total weight of the oxygen atoms will be 48 (i.e., 3×16) microcriths. Let x be the molecular weight of potassium chlorate, and make the proportion $x : 48 :: \text{wt. of potassium chlorate taken} : \text{wt. of oxygen lost}$. Solving the proportion for x , you should obtain somewhere near the number 122.

EXPERIMENT 19

Molecular Weight of Potassium Chlorid

Knowing the molecular weight of potassium chlorate, it is easy for us to find the molecular weight of potassium chlorid. From Experiment 6 and 25 c, Part I., we know that the residue left in the crucible after heating the chlorate is potassium chlorid; hence the molecular weight of potassium chlorate minus 48 is the molecular weight of potassium chlorid.

By similar methods the molecular weights of a large number of compounds that cannot be volatilized have been determined. But it must be remembered that we cannot be sure that the numbers we obtain by the chemical method are the correct ones, until we have studied a great number of chemical changes into which the elements that constitute the compounds enter. Besides this, we must verify these numbers by means of other facts.

Atomic Weights in General. By comparing all the various numbers obtained from experiments for determining chemical, electrical, isomorphical, and thermal equivalents, and molecular weights obtained by the physical and chemical

methods, chemists have obtained a series of numbers which represent so accurately the atomic weights of the elements that there is practically no doubt of their truth within very small prescribed limits. These numbers are used without showing appreciable error in numberless chemical investigations, and they are to-day accepted to be as near the actual numbers as our present methods of analysis admit.

Periodicity of the Elements. If we write horizontally in order, in seven columns, the numbers representing the atomic weights of the elements (omitting hydrogen), we shall see that those in the same vertical column resemble each other to a remarkable degree both in their chemical properties and in the similarity of their compounds. A short study of this table will be profitable.

In the first column, we find first the strong alkalis; in the second, the closely allied elements calcium, barium, and strontium, and also the natural group magnesium, zinc, and cadmium. In the third are found most of the rarer elements; in the fifth, we have the similar elements oxygen, sulfur, selenium, and tellurium; and in the seventh, we have the halogens. Thus it seems that the properties of the elements depend upon their atomic weights, with similar properties recurring at regular intervals. Besides the regular column, an extra column is added for those elements which for some unknown reason do not seem to fit in with the regular groups. It will be noticed that, in the irregular groups, the elements have almost the same atomic weights and similar properties. Vacant spaces have been left, as there are at present no elements known whose atomic weights fit in those spaces. When the table was first made out, Mendelejeff, to whom we are indebted for this remarkable discovery, left two vacant spaces between the elements zinc and arsenic. Mendelejeff predicted the properties of both these elements,

	I.	II.	III.	IV.	V.	VI.	VII.	VIII.			
1	H 1										
2	Li 6.97	Be 9.01	B 10.86	C 11.91	N 13.93	O 15.88	F 18.91				
3	Na 22.88	Mg 24.1	Al 26.91	Si 28.18	P 30.79	S 31.83	Cl 35.18				
4	K 38.82	Ca 39.76	Sc 43.78	Ti 47.79	V 50.99	Cr 51.74	Mn 54.57	Fe 55.6	Ni 58.24	Co 58.55	Cu 63.12
5	(Cu) 63.12	Zn 64.91	Ga 69.38	Ge 71.93	As 74.44	Se 78.58	Br 79.34				
6	Rb 84.78	Sr 86.95	Y 88.35	Zr 89.72	Cb 93.02	Mo 95.26	—	Ru 100.91	Rh 102.23	Pd 105.56	Ag 107.11
7	(Ag) 107.11	Cd 111.54	In 112.99	Sn 118.15	Sb 119.52	Te 126.52	I 125.89				
8	Cs 131.89	Ba 136.39	La 137.59	Ce 138.90	Pr 139.41	Nd 142.52	—	—	—	—	—
9	(—)	—	—	Gd 155.57	—	—	—	—	—	—	—
10	Tb 158.80	Er 165.06	Tu 169.4	Yb 171.88	Ta 181.45	W 183.43	—	Os 189.55	Ir 191.66	Pt 193.41	Au 195.74
11	(Au) 195.74	Hg 198.49	Tl 202.61	Pb 205.36	Bi 206.54	—	—	—	—	—	—
12	—	—	—	Th 230.87	—	U 237.77	—	—	—	—	—

which have since been discovered and found to agree in a wonderful degree with what had been expected.

This table has been of service in determining what numbers should be taken as the atomic weights of certain elements. For instance, the atomic weight of uranium was supposed to be 120, but that number threw it out of place in the table. The number was then doubled, which placed it in the position in the table where from its properties it evidently belonged. This weight has also been verified by its specific heat determination.

Prout's Hypothesis. In 1815, an English chemist, Prout, advanced the theory that, inasmuch as the atomic weights of the elements were simple multiples of the atomic weight of hydrogen, therefore they were all composed of definite portions of one kind of matter, and that the quantity of this kind of matter in an atom determined its properties. However attractive this theory may be, many of the greatest investigators believe it to be untenable from the fact that, as accuracy in the determination of atomic weights has increased, it has been found that comparatively few of them can be expressed by whole numbers.

Symbols. We are now ready for some simple method of representing chemical changes, not only qualitatively as in Part I., but quantitatively also. Chemists have agreed to represent the atom of each element by a letter, which shall stand both for the atom itself and also for its atomic weight. Thus the letter H stands for an atom of hydrogen, and also for its atomic weight 1; the letter O stands for an atom of oxygen, and also for its atomic weight 16. By combining these letters, we may represent molecules. Thus H_2SO_4 signifies a molecule of sulfuric acid, molecular weight 98; the small numerals mean that there are two atoms of hydrogen and four atoms of oxygen, besides the single atom of sulfur.

Aluminum	Al	26.91	Chlorine	Cl	35.18	Hydrogen	H	1.00	Neodymium	Nd	142.52	Ruthenium	Ru	100.91	Thallium	Tl	202.61
Antimony	Sb	119.52	Chromium	Cr	51.74	Indium	In	112.99	Nickel	Ni	58.24	Samarium	Sm	149.13	Thorium	Th	230.87
Arsenic	As	74.44	Cobalt	Co	58.55	Iodine	I	125.89	Nitrogen	N	13.93	Scandium	Sc	43.78	Thulium	Tu	169.40
Barium	Ba	136.39	Columbium	Cb	93.02	Iridium	Ir	191.66	Osmium	Os	189.55	Selenium	Se	78.58	Tin	Sn	118.15
Bismuth	Bi	206.54	Copper	Cu	63.12	Iron	Fe	55.60	Oxygen	O	15.88	Silicon	Si	28.18	Titanium	Ti	47.79
Boron	B	10.86	Erbium	Er	165.06	Lanthanum	La	137.59	Palladium	Pd	105.56	Silver	Ag	107.11	Tungsten	W	183.43
Bromine	Br	79.34	Fluorine	F	18.91	Lead	Pb	205.36	Phosphorus	P	30.79	Sodium	Na	22.88	Uranium	U	237.77
Cadmium	Cd	111.54	Gadolinium	Gd	155.57	Lithium	Li	6.97	Platinum	Pt	193.41	Strontium	Sr	86.95	Vanadium	V	50.99
Cæsium	Cs	131.89	Gallium	Ga	69.38	Magnesium	Mg	24.10	Potassium	K	38.82	Sulfur	S	31.83	Ytterbium	Yb	171.88
Calcium	Ca	39.76	Germanium	Ge	71.93	Manganese	Mn	54.57	Praseodymium	Pr	139.41	Tantalum	Ta	181.45	Yttrium	Yt	88.35
Carbon	C	11.91	Glucinum	Gl	9.01	Mercury	Hg	198.43	Rhodium	Rh	102.23	Tellurium	Te	126.52	Zinc	Zn	64.91
Cerium	Ce	138.30	Gold	Au	195.74	Molybdenum	Mo	95.26	Rubidium	Rb	84.78	Terbium	Tb	158.80	Zirconium	Zr	89.72

The table on page 128 gives a list of the elements with their symbols and atomic weights.*

Determination of Molecular Formulae. How do we know that sulfuric acid is H_2SO_4 , that copper sulfate is CuSO_4 , that alcohol is $\text{C}_2\text{H}_6\text{O}$, etc.?

First we must determine by analysis the percentage composition of the compound. Let us take as an example the last one above mentioned, alcohol. When analyzed, alcohol is found to contain

Carbon	52.17 %
Hydrogen	13.05 %
Oxygen	34.78 %
	<hr/> 100.00 %

We find that the ratios 52.17 : 13.05 : 34.78 can be expressed in any of the following ways:—

	4 : 1 : 2.66
	8 : 2 : 5.32
	12 : 3 : 7.98
	16 : 4 : 10.64
	20 : 5 : 13.30
and	24 : 6 : 15.96

the last of which is practically 24 : 6 : 16. Therefore 24 grams of carbon, 6 grams of hydrogen, and 16 grams of oxygen unite, forming 46 grams of alcohol. Or, expressing the quantity of each element in microcriths, 24 microcriths of carbon, 6 microcriths of hydrogen, and 16 microcriths of oxygen unite, forming 46 microcriths of alcohol. But, since an atom of carbon weighs 12 microcriths, an atom of hydrogen 1 microcrith, and an atom of oxygen 16 microcriths, we see at once that the simplest composition of a molecule of alcohol would be two atoms of carbon, six atoms of hydrogen, and one atom of oxygen, i.e., $\text{C}_2\text{H}_6\text{O}$.

* From report of American Chemical Society, 1899.

We might just as well use the ratios

$$\begin{array}{l} 48 : 12 : 32 \\ \text{or} \quad 64 : 18 : 48 \end{array}$$

We should then have obtained either the formula $\text{C}_4\text{H}_{12}\text{O}_2$ or $\text{C}_6\text{H}_{18}\text{O}_3$. The only way we have of deciding which is correct is by comparing the molecular weights which these formulae give, with the density of alcohol in the state of vapor.

$\text{C}_2\text{H}_6\text{O}$ gives a molecular weight of $24 + 6 + 16 = 46$

$\text{C}_4\text{H}_{12}\text{O}_2$ gives a molecular weight of $48 + 12 + 32 = 92$

$\text{C}_6\text{H}_{18}\text{O}_3$ gives a molecular weight of $72 + 18 + 48 = 138$

By weighing alcohol in the state of vapor and comparing it with hydrogen, we found in Experiment 12, Part II. that its destiny is 23. Therefore its molecular weight must be 2×23 , or 46. Hence we take the formula $\text{C}_2\text{H}_6\text{O}$ as the correct one. In the case of compounds whose density cannot be found, we take the simplest formula.

The simplest way numerically to find the formula when the percentage composition is given is as follows :

$$\begin{array}{lcl} \text{Carbon} & 52.17 \div 12 = 4.34 & \\ \text{Hydrogen} & 13.05 \div 1 = 13.05 & \\ \text{Oxygen} & 34.78 \div 16 = 2.17 & \end{array} \left. \vphantom{\begin{array}{l} 52.17 \\ 13.05 \\ 34.78 \end{array}} \right\} \begin{array}{l} 2 \text{ C} \\ 6 \text{ H} \\ 1 \text{ O} \end{array}$$

Rule. Divide the percentage of each element by its atomic weight. Find what the resulting ratios become, if the smallest number be taken as unity. If there is only one atom of that element, the ratios will then be expressed by whole numbers. If the resulting ratios are not whole numbers, try successively what the ratios would become, if the smallest number were taken as 2, 3, etc.

Examples. Deduce the formulae for the following substances : —

1. {	Hydrogen	11.12	7. {	Copper	39.62
	Oxygen	88.88		Sulfur	20.13
		100.00		Oxygen	40.25
	Density, 9				100.00
2. {	Hydrogen	5.88	8. {	Hydrogen	3.06
	Oxygen	94.12		Phosphorus	31.64
		100.00		Oxygen	65.30
3. {	Manganese	72.05			100.00
	Oxygen	27.95			
		100.00			
4. {	Silicon	46.67	9. {	Potassium	45.95
	Oxygen	53.33		Nitrogen	16.45
		100.00		Oxygen	37.60
5. {	Iron	70.01			100.00
	Oxygen	29.99			
		100.00			
6. {	Carbon	92.3	10. {	Nitrogen	29.17
	Hydrogen	7.7		Hydrogen	8.33
		100.00		Carbon	12.50
	Density, 13.			Oxygen	50.00
					100.00

Valence. Let us suppose that we know the molecular formulae for all the known compounds.

Then, if we arrange a few of the simplest thus : —

I.	II.	III.	IV.
H F	H ₂ O	H ₃ N	H ₄ C
H Cl	H ₂ S	H ₃ P	H ₄ Si
H Br	H ₂ Se	H ₃ As	—
H I	H ₂ Te	H ₃ Sb	—

we shall see that an atom of each element possesses the power of uniting with a certain number of hydrogen atoms.

Again, let us make a table thus : —

Acids	Salts		
H Cl	Ag Cl	Cu Cl ₂	Al Cl ₃
H N O ₃	Ag N O ₃	Cu (N O ₃) ₂	Al (N O ₃) ₃
H ₂ S O ₄	Ag ₂ S O ₄	Cu S O ₄	Al ₂ (S O ₄) ₃
H ₃ P O ₄	Ag ₃ P O ₄	Cu ₃ (P O ₄) ₂	Al P O ₄

We here see that an atom of each element possesses the power of replacing a certain number of hydrogen atoms in a compound. This property of an atom is called its *valence*. We might define *valence as the quantity of combining power an atom has compared with that of an atom of hydrogen*. Atoms that combine with or replace an atom of hydrogen are called monovalent, or monads. Those which combine with or replace two atoms of hydrogen are called divalent, or dyads. There are also trivalent, tetravalent, etc., atoms. For some of the elements, the number representing the valence varies. For instance, nitrogen has a variable valence, being sometimes triad and sometimes pentad.

The following table shows the valence of the more common elements.

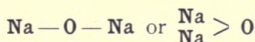
Al 3	Cr 3	Na 1
As 3, 5	Cu 2	Ni 2
Ag 1	F 1	O 2
B 3	Fe 2, 4	P 3, 5
Ba 2	H 1	Pt 4
Bi 3, 5	Hg 2	Pb 2
Br 1 (3, 5, 7)	I 1 (5)	S 2, 4, 6
C 4 (2)	K 1	Si 4
Ca 2	Li 1	Sn 2, 4
Cl 1 (3, 5, 7)	Mg 2	Sr 2
Cd 2	Mn 2	Sb 3, 5
Co 2	N 3, 5	Zn 2

Graphic Formulae. Although we do not know the nature of chemical affinity, still, as we have seen, we can express it quantitatively. We can represent monad, dyad, triad, etc., valence by lines; a monad element being written H —, a

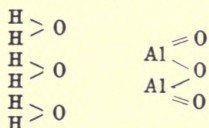
dyad, — O —, a triad, $\begin{array}{c} | \\ \text{B} - \\ | \end{array}$, and so on.

Let us start with the compound H₂O, and write it H — O — H,

or $\begin{array}{c} \text{H} \\ \diagup \\ \text{H} \end{array} > \text{O}$, understanding by this that one atom of oxygen with a valence of two is united with two atoms of hydrogen, each having a valence of one. Suppose now that we replace the monad hydrogen atoms by two monad sodium atoms. We shall then have for sodium oxid the formula

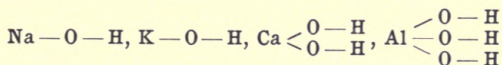


In like manner for potassium oxid, we should have $\text{K} - \text{O} - \text{K}$. If we replace the two monad hydrogen atoms by a single dyad atom, such as calcium, we should have $\text{Ca} = \text{O}$ for calcium oxid. Using the triad aluminum atom, we should have to replace the six monad hydrogen atoms of three molecules of water by two aluminum atoms.



In like manner other graphic formulae of other oxids may be written.

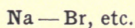
The corresponding hydroxids, of course, will be



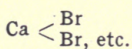
the metals replacing only one hydrogen atom from each water molecule.

In the case of an acid, we write $\text{H} - \text{Cl}$ for hydrochloric, $\text{H} - \text{Br}$ for hydrobromic, etc. We then have:

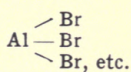
SODIUM SALTS



CALCIUM SALTS

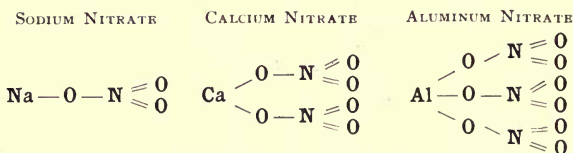


ALUMINUM SALTS

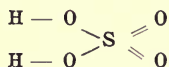


We have seen that all acids contain hydrogen, but not necessarily oxygen. The greater number of acids, however, contain oxygen, and it has been shown by a number of investigators that at least part of the oxygen in these acid molecules is closely associated with the replaceable hydrogen. Thus in nitric, HNO_3 , we have the graphic formula $\text{H} - \text{O} - \text{N} \begin{smallmatrix} \text{O} \\ \parallel \\ \text{O} \end{smallmatrix}$. This group, $-\text{O} - \text{H}$, which appears also in the bases $\text{Na} - \text{O} - \text{H}$, $\text{K} - \text{O} - \text{H}$, etc., is called the *hydroxyl* group, and is evidently a monad. We know that sodium nitrate differs from nitric acid, in that it contains an atom of sodium in place of an atom of hydrogen.

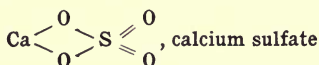
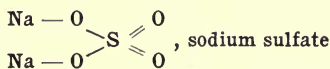
We have, therefore, the following formulæ:

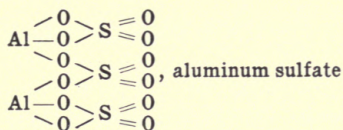


In the same way, it is believed, sulfuric acid, H_2SO_4 , contains two hydroxyl groups, thus: —

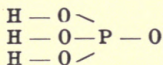


From this we may obtain the salts

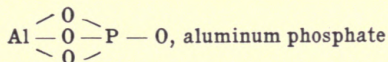
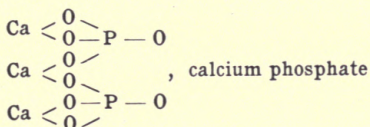
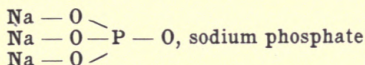




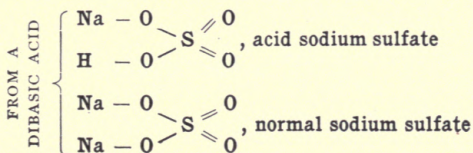
Similarly phosphoric acid, H_3PO_4 , contains three hydroxyl groups, thus : —

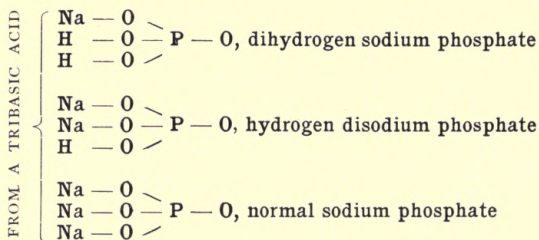


from which we have the salts,



Acids that contain one hydroxyl group are called *monobasic*, those containing two such groups are called *dibasic*, those containing three are called *tribasic*, and so on. When all the hydrogen of the hydroxyl groups is replaced by a metal, the resulting salt is said to be *normal*. Evidently it is possible that only the hydrogen of one hydroxyl group in a dibasic or higher acid may be replaced, thus giving rise to *acid* salts. Hence two different salts containing the same metal may be obtained from a dibasic acid, three from a tribasic acid, and so on. We may illustrate this as follows :—

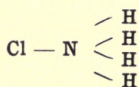




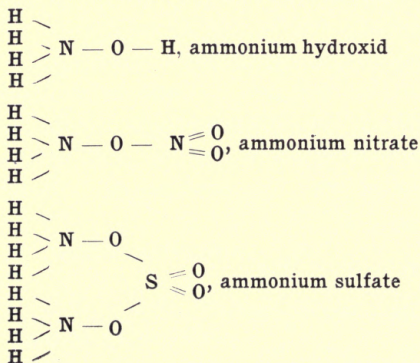
At present little is known concerning the molecular structure of basic salts. Some think that to form these salts, additional molecules of the base replace part of the oxygen of the acid; while others think that they are combinations of the normal salt molecules with extra base molecules.

The molecular formula for ammonia is NH_3 , or $\text{N} \begin{array}{l} \diagup \text{H} \\ - \text{H} \\ \diagdown \text{H} \end{array}$

We have formed ammonium chlorid, NH_4Cl , synthetically by the union of ammonia and hydrochloric acid, giving of course



and showing the pentad valence of nitrogen. The group NH_4 is evidently a monad group, and as such appears in



This group is called the *ammonium group*, and acts very much like a metal.

Thus the student will see the possibility of representing through the eye relations that would otherwise be incompletely realized. It must not, however, be thought that these structural formulae are intended to represent the actual position of the atoms of a molecule with respect to each other. The intention is merely to emphasize certain relations that experiment and reason have shown must exist between the atoms in the molecule.

Examples. Write the graphic formulae for silver nitrate, aluminum hydroxid, calcium carbonate, and ammonium phosphate.

Positive and Negative Elements. Chemical compounds can be decomposed by electricity. In all such decompositions, part of the atoms appear at the negative pole, and part at the positive pole. Those which appear at the negative pole are called electro-positive, or simply positive, elements; and those which appear at the positive pole are called electro-negative, or simply negative. One element, however, need not always appear at the same pole. When liberated from its union with one element, it may appear at the negative pole; while, when liberated from another, it may appear at the positive pole. The non-metals are negative with respect to the metals, but it is readily seen that, compared with one another, they may be one or the other according to what elements they are. The strength of chemical combination depends upon this, the rule being that the more electrically remote the elements are, the stronger their union.

Naming of Compounds. *Binary* compounds are those composed of two elements. All others are called *ternary*. The names of binary compounds, the names of ternary com-

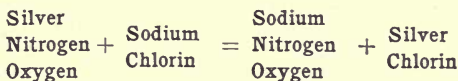
pounds that do not contain oxygen, and the names of compounds of positive elements with hydrogen and oxygen, end in *id*. Examples respectively, NaCl, sodium chlorid; KCN, potassium cyanid; $\text{Zn}(\text{OH})_2$, zinc hydroxid. In all cases the positive element is named first; then the negative element or radical, with the suffix. When more than one compound is formed between the same elements, the name of the positive element ends in *ous* or *ic*, according as there is less or more of the negative element or radical. Examples HgI , mercurous iodid; HgI_2 , mercuric iodid.

An acid takes its name from the characteristic element in it; and when there is more than one acid from the same element, the name ends in *ous* or *ic* according as there is less or more oxygen. Examples, H_2SO_3 , sulfurous acid; H_2SO_4 , sulfuric acid. If there are more than two, the prefixes *hypo*, *meta*, *per*, and *pyro*, are used also. Examples H_2SO_2 , *hypo*-sulfurous acid; HClO_4 , *perchloric* acid. The names of acids that contain no oxygen have the prefix *hydro*. Example, HCl , *hydrochloric* acid.

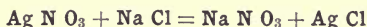
Salts take their names from the name of the positive element and that of the acid. The positive element is named first, and then the acid ending in *ite* or *ate*. The names of salts from "*ous*" acids end in *ite*; those from "*ic*" acids end in *ate*. Examples, Na_2SO_3 sodium sulfite; Na_2SO_4 , sodium sulfate. When there is more than one salt from the same element and acid, the name of the positive element ends in *ous* or *ic*, according as there is less or more of the acid radical. Examples, HgNO_3 mercurous nitrate; $\text{Hg}(\text{NO}_3)_2$, mercuric nitrate.

Writing of Reactions. In order to be able to write reactions by means of chemical symbols, it will be necessary for the student to learn the valence of each of the elements.

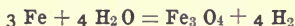
Suppose we wish to express qualitatively the fact that the action of silver nitrate upon sodium chlorid gives silver chlorid and sodium nitrate, we may do so with a formula like that used in Part I.



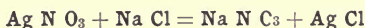
In order to express it quantitatively, we must make use of the molecular formulae of the molecules that act upon each other, keeping in mind the valence of the atoms of which these molecules are composed. Thus,



Sometimes it will be necessary to use more than one molecule of each substance, as in the case of the action of steam upon red-hot iron.

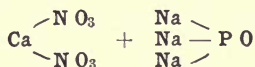


These formulae not only show what action has taken place between the molecules, but also how much matter is involved. Thus they indicate the fact that there is just as much matter after the action as there was before. For instance, the formula,

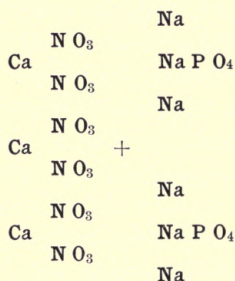


stands for the following sentence. One molecule of silver nitrate, molecular wt. 168.68, acts with one molecule of sodium chlorid, molecular wt. 58.06, producing one molecule of sodium nitrate, molecular wt. 84.45, and one molecule of silver chlorid, molecular wt. 142.29.

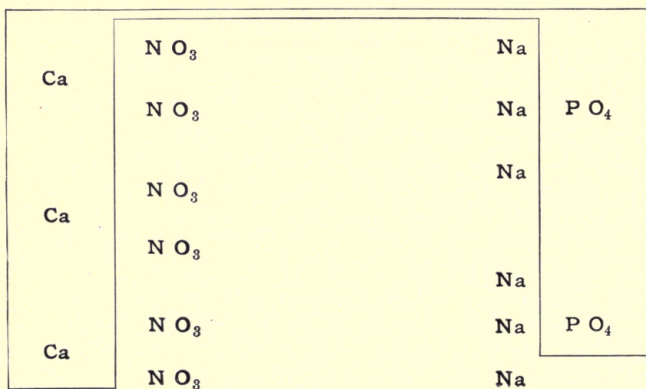
The student will be aided considerably in writing reactions by the following illustration. Suppose we wish to write the reaction for calcium nitrate and sodium phosphate. Write the formulae for the two compounds thus:—



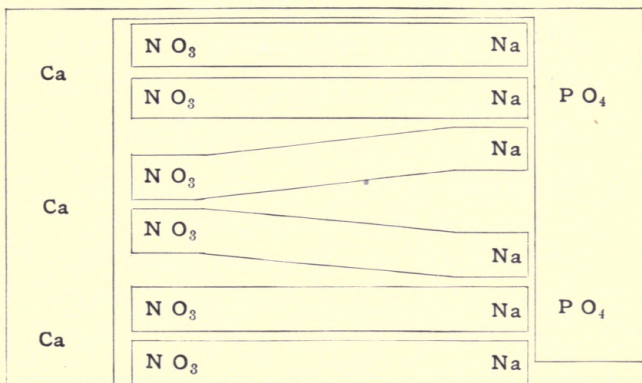
Knowing that calcium is dyad in valence, and that phosphoric acid is a tri-basic acid, we see that we must take enough molecules of calcium nitrate to make the total valence of the calcium atoms equal to 6, and enough sodium phosphate molecules to make the total valence of the sodium atoms equal to 6, i.e., in both cases the least common multiple of 2 and 3, the valence of calcium and the basicity of phosphoric acid respectively. We then have



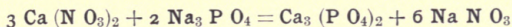
The three Ca atoms will now replace the six sodium atoms, unite with the two PO_4 groups, and give one molecule of calcium phosphate, indicated thus:—



The remaining Na's and NO_3 's will unite, forming six molecules of sodium nitrate, thus : —



We then have the complete equation



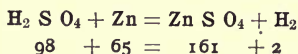
EXPERIMENT 20

Write the reactions for all the experiments in Part I., using chemical symbols.

Stoichiometry. Since chemical action takes place between molecules of elements or compounds, in order to find the mass weights of elements or compounds formed, knowing the mass weights of elements or compounds used, we have only to make use of the following: — *First write the reaction representing the chemical change. Then make the proportion, — molecular weight of the given substance is to the molecular weight of the required substance as the mass weight of the given substance is to the mass weight of the required substance.*

Let us take an example.

How many grams of hydrogen will be evolved, when 10 grams of zinc are dissolved in sulfuric acid?



In this case we have

$$65 : 2 :: 10 : x. \quad \therefore x = .30 + \text{ grms.}$$

Suppose we wish to know the number of grams of zinc sulfate formed. We then have

$$65 : 161 :: 10 : x. \quad \therefore x = 24.7 + \text{ grms.}$$

Suppose we wish to know the number of grams of sulfuric acid required to dissolve 10 grams of zinc. We then have

$$65 : 98 :: 10 : x. \quad \therefore x = 15.0 + \text{ grms.}$$

In case we wish to know what the volume of the hydrogen evolved would be at 0° and 760 mm., all we have to do is to divide the weight of the hydrogen by .0896, the weight of a liter of hydrogen at 0° and 760 mm. (See Exp. 10, Part II.)

If we wish to know what this volume would be at any required temperature and pressure, we can easily find it by using the formula

$$\frac{V P}{273 + t} = \frac{V' P'}{273 + t'} \quad (\text{See Exp. 6 and 7, Part II.})$$

Examples. 1. How many grams of potassium chlorate must be used to obtain 100 grams of oxygen? *Ans.* 264 + grms.

2. In order to fill a balloon, 150 kilograms of hydrogen are necessary. How much zinc and sulfuric acid will be required to produce the gas?

Ans. Zinc, 4875 kilos.
Sulfuric acid, 7350 kilos.

3. What would the volume of the gas be at 0° and 760 mm. pressure? What would it be on a day when the temperature was 20° and the pressure 755 mm.?

Ans. 1666.6 + liters.
1800.5 + liters.

4. How many grams of silver nitrate would be required to make 20 grams of silver chlorid? *Ans.* 23.7 + grms.

5. How many grams of iron sulfid must be used to produce, at 0° and 760 mm., 10 liters of hydrogen sulfid, 1 liter of hydrogen sulfid weighing 1.52 grams? *Ans.* 39.3 + grms.

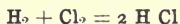
6. Calculate the amount of manganese dioxid that must be used to produce, at 0° and 760 mm., 10 liters of chlorin from hydrochloric acid, 1 liter of chlorin weighing 3.17 grms.

Ans. 38.9 grms.

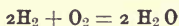
7. What weight of copper would be used in making 20 grams of copper nitrate by dissolving the copper in nitric acid?

Ans. 6.9 grms.

The reactions also indicate the relative volumes involved in the case of gaseous factors and products. By the equation



we indicate that 1 volume of hydrogen combines with 1 volume of chlorin, forming 2 volumes of hydrochloric acid. In the same way



indicates that 2 volumes of hydrogen unite with 1 volume of oxygen, forming 2 volumes of water vapor.

Examples. 1. 18 cc. of hydrogen are mixed with 10 cc. of chlorin and exploded. What gases are formed, and what are their volumes?

2. If 2 volumes of nitric oxid and 5 volumes of hydrogen are united, what volume of ammonia is produced?

3. To a certain volume of hydrogen sulfid gas, was added

125 cc. of chlorine gas, which was entirely consumed. What volume of hydrochloric acid gas was evolved?

4. If 150 liters of marsh gas be exploded with 300 liters of oxygen, what volume of carbon dioxide would result?

Calculation of Percentage Composition, having given the Molecular Formula. It is often required that the student should be able to calculate the percentage composition of a compound when the molecular formula is known. Suppose we have given the formula for alcohol, which is C_2H_6O .

Carbon	$2 \times 12 = 24$
Hydrogen	$6 \times 1 = 6$
Oxygen	$1 \times 16 = 16$
Molecular weight	$= 46$

In alcohol, therefore, there are 12 parts by weight of carbon, 6 parts of hydrogen, and 16 of oxygen. This reduced to the decimal system becomes

Carbon	$\frac{24}{46} = 52.17\%$
Hydrogen	$\frac{6}{46} = 13.05\%$
Oxygen	$\frac{16}{46} = 34.78\%$
	<hr/> 100.00%

In case there is water of crystallization in the compound, the water molecules combined with each molecule of the compound must also be taken into account.

Examples. 1. Calculate the percentage composition of $AgCl$; $NaNO_3$; $Ca_3(PO_4)_2$; $CHCl_3$; $K_4Fe(CN)_6$; $(NH_4)_2SO_4$.

2. Calculate the percentage composition of $MgSO_4 \cdot 7H_2O$; $HNaNH_4PO_4 \cdot 4H_2O$.

Thermochemistry. It is necessary even for the student of elementary chemistry to understand at least the fundamental

relations between heat and chemical change. All chemical changes are accompanied either by the using up or the giving out of heat.* For instance, when 2 grams of hydrogen unite with 18 grams of oxygen, 68924 heat units are liberated. Such actions are called *exothermic*. There are also chemical changes which absorb heat. For instance, to unite 1 gram of hydrogen with 127 grams of iodine to form hydriodic acid, requires 6000 heat units. Such actions are called *endothermic*.

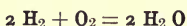
* It is well to note here the relation between chemical energy and other forms of energy. *Energy* is the power that matter has of doing *mechanical work*, i.e., of overcoming resistance. Energy may be of two kinds, either *kinetic* or *potential*. *Kinetic energy* is the energy matter has by virtue of its motion, while *potential energy* is the energy it has by virtue of its position or condition. For instance, a flying cannon ball can overcome resistance by reason of its motion. On the other hand, the same cannon ball, supported at a height from the ground, has in it, by virtue of its position, the power of acquiring motion when the support is removed, and of thus overcoming resistance. The chemical energy possessed by matter is potential energy. Coal, on account of the affinity of carbon for oxygen, will burn, and so give out heat. This heat in turn may be utilized to boil water, the steam from which in expanding will move the piston of an engine, and thus do mechanical work. Also the chemical energy of matter may be transformed into electric energy as in the galvanic cell. This electrical energy may be utilized to drive a motor, and thus do mechanical work.

It has been shown that carbon dioxide neither burns nor supports combustion (see Exp. 86). The reason for this is that the carbon has already united with all the oxygen that it can hold; therefore no further combustion can take place. Just as the cannon ball would have to be lifted again to possess potential energy, so the carbon dioxide would have to be decomposed in order to possess chemical energy again. Carbon dioxide does not support combustion, because the affinity between carbon and oxygen is stronger than that between oxygen and most other elements. However, if an element having a stronger affinity for oxygen, as for instance potassium at a high temperature, is brought in contact with carbon dioxide, then the gas supports combustion.

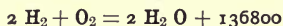
We shall use as our unit of heat the *calorie*, i.e., the amount of heat necessary to raise one gram of water one degree. The number of calories given out by the burning of one gram of various substances has been determined by a number of investigators. The following table gives a few examples.

Hydrogen	34180	Sulfur	2220
Carbon	8080	Zinc	1300
Phosphorus	5747	Iron	1181

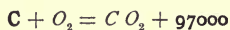
We have learned that the equation



is in itself a statement of the law of the indestructibility of matter. But we know that, in the burning of hydrogen, heat is evolved. This equation, as it stands, tells us nothing in regard to this fact. It is very easy, however, to supplement the equation so that it shall convey the full meaning. Let the chemical symbols in an equation stand for the number of grams corresponding to its molecular weight. Thus O_2 stands for 32 grams of oxygen, HCl stands for 35.2 grams of hydrochloric acid. By adding to the chemical equation the number of calories of heat evolved as the result of the action, we express the meaning in full. In the example of burning hydrogen we have



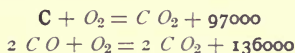
which signifies that 4 grams of hydrogen on uniting with 32 grams of oxygen form 36 grams of water, and at the same time give out 136800 calories of heat. In the case of endothermic reactions, the notation is the same, except that the — sign instead of the + sign is used. In addition to this, heavy type is used to represent solids, ordinary type liquids, and italics, gases. Thus,



- signifies that 12 grams of solid carbon burned in 32 grams of gaseous oxygen gives 44 grams of gaseous carbon dioxide together with 97000 calories of heat.

In a great many cases where the heat liberated cannot be found by experiment, it can be found by calculation.

Example. Required to find the heat liberated when carbon is oxidized to carbon monoxide. It is found that when 12 grams of carbon is oxidized to carbon dioxide, 97000* calories are liberated, and that when 28 grams of carbon monoxide is oxidized to carbon dioxide, 136000 † calories are liberated. We then have



By dividing the second equation by 2, we obtain 68000 cal. as the amount of heat liberated when 44 grams of carbon dioxide are formed. We divide by 2 in order to obtain the same weight of carbon dioxide as in the first equation. Subtracting 68000 from 97000, we evidently obtain the number of calories given out when carbon is oxidized to the monoxide, i.e., 29000 cal. We therefore have the equation



EXPERIMENT 21

Heat of Chemical Action

Dilute 50 cc. of concentrated sulfuric acid with 250 cc. of water, and allow it to cool. Weigh a beaker of thin glass large enough to hold the acid, and place it in a larger beaker. Pack wool around and under the inner beaker in such a way that the inner beaker can be easily removed. Let this inner beaker be used as a calorimeter. Pour the acid into the

* More correctly 96960.

† More correctly 135920.

calorimeter. Clean a piece of sheet zinc about $2\frac{1}{2}$ inches wide by 5 long. Take the temperature of the acid, then plunge the zinc into it. When the temperature has risen four or five degrees, remove the zinc, stir, and take the temperature carefully. Wash the zinc clean, dry it, and weigh again.

It will be necessary to know the specific heat of the solution after the zinc has been acted upon. To find this, perform an experiment exactly like Experiment 15, Part II. with the exception of using this calorimeter and liquid just as they are, instead of the lemonade shaker and water. Calculate the specific heat of the liquid, knowing the specific heat of lead to be .031.

Arrange your calculations as follows :—

To find the specific heat of the liquid.

Wt. of liquid = 1 =

Wt. of calorimeter = c =

Wt. of shot = s =

Specific heat of shot = .031

Specific heat of glass = .2

Temperature liquid = t =

Temperature liquid and shot = t' =

Temperature of shot = t'' =

Let x = specific heat of the liquid.

$$.031 s (t'' - t') = .2 c (t' - t) + x l (t' - t)$$

Solve for x .

To find the amount of heat liberated when one gram of zinc is dissolved in sulfuric acid.

Wt. of zinc dissolved = z =

Wt. of water and acid = 1 =

Wt. of calorimeter = c =

Specific heat of glass = .2

Specific heat of liquid = s (found in foregoing calculation)

Temperature before action = t =

Temperature after action = t' =

Let x = the number of calories liberated when one grm. of zinc is dissolved.

The amount of heat evolved by the action of the acid on the zinc equals the amount of heat gained by the solution and the calorimeter.

$$x z = s l (t' - t) + .2 c (t' - t)$$

Solve for x and obtain the number of calories liberated when one gram of zinc is dissolved. Find $65x$ in conformity with principle laid down on page 146.

Write the reaction for the chemical change together with the heat evolved.

EXPERIMENT 22

Heat of Neutralization

Weigh out about 25 grams of c.p. concentrated sulfuric acid. Dilute this with 200 cc. of water, and allow it to cool. Calculate (see page 142) the number of grams of sodium hydroxid necessary to neutralize the acid, take about one-sixth more than this amount, dissolve it in 200 cc. of water, and allow the solution to cool. Let the two vessels containing respectively the acid and alkali stand side by side until they are of the same temperature. Then pour them together in the glass calorimeter used in Exp. 21, and note the rise in temperature.

Arrange your calculations as follows:—

Wt. of sulfuric acid = s =

Wt. of sodium hydroxid = h =

Wt. of calorimeter = c =

Specific heat of glass = $.2$

Temperature before mixing = t =

Temperature after mixing = t' =

Wt. of water = w =

Specific heat of solution = a =

Find a , the specific heat of the sodium sulfate solution, in the same way as you did that of zinc sulfate in Experiment 21. The heat gained by the solution and calorimeter equals the heat evolved by neutralization.

$$\text{Heat gained by solution} = (s + w + h) (t' - t) a$$

$$\text{Heat gained by calorimeter} = .2 c (t' - t)$$

$$\text{Heat of neutralization} = sx$$

Therefore

$$(s + w + h) (t' - t) a + .2 c (t' - t) = sx$$

$$x =$$

$$98 x =$$

We find $98x$ in conformity with the principle laid down on page 146.

Write the reaction for the chemical change together with the heat evolved.

EXPERIMENT 23

Heat of Solution and of Hydration

a . Place in a weighed calorimeter about 350 grams of water weighed accurately to one gram. Powder exactly 40 grams of anhydrous magnesium sulfate, place it in a beaker, and cover it with a watch glass. Allow the calorimeter, containing the water, and the beaker to stand side by side until they are of the same temperature. Then pour the powdered salt into the water, and stir the solution with a thermometer, taking the temperature when all is dissolved.

Arrange your calculations as follows:—

$$\text{Wt. of magnesium sulfate} = m =$$

$$\text{Wt. of calorimeter} = c =$$

$$\text{Wt. of water} = w =$$

$$\text{Temperature before mixing} = t =$$

$$\text{Temperature after mixing} = t' =$$

$$\text{Specific heat of solution} = a =$$

$$\text{Specific heat of glass} = .2$$

Find a in the same way as in Experiments 21 and 22.

The heat here given out is made up of two different heats; first, the heat given out by the anhydrous salt taking on water of crystallization; second, the heat used up by the salt dissolving in water.

We then have the equation:—

Heat gained by the solution and calorimeter = heat of hydration and of solution.

$$\text{Heat gained by the solution} = (w + m) (t' - t) a$$

$$\text{Heat gained by the calorimeter} = .2 c (t' - t)$$

$$\text{Heat of hydration and of solution} = mx$$

Therefore

$$(w + m) (t' - t) a + .2 c (t' - t) = mx$$

$$x =$$

$$120 x =$$

The number 120 is taken for the same reason that 98 was taken in the last experiment. (See page 146.)

You should obtain as a result somewhere near the number 20280.

b. Now repeat the operation, using exactly 82 grams of the salt with its water of crystallization ($\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$). In this case, calculate the heat lost in dissolving 246 grams of magnesium sulfate. You should obtain a number somewhere near 3800.

We see then that the heat of hydration (a positive heat) must be 3800 greater than the heat of solution (a negative heat), i.e., 24080.

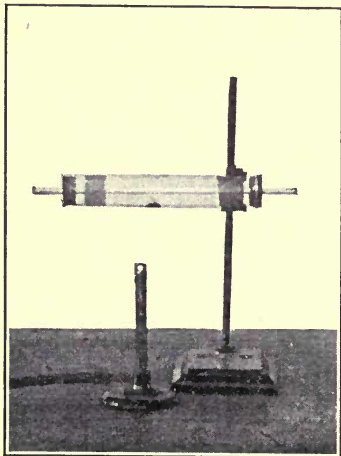
Dissociation. We have found (Exp. 42, Part I.) that when ammonium chlorid is heated it passes directly from the solid to the gaseous state. If the density of ammonium chlorid is found in the state of vapor, the number obtained is 13.34. Now, according to the hypothesis of Avogadro, the number

should be 26.69, since the molecular weight is 53.38. The densities of hydrochloric acid gas and ammonia are, however, 18.18 and 8.5 respectively; which, if added, give 26.68, showing that undoubtedly, when ammonium chlorid is vaporized, it is not made up of molecules of ammonium chlorid, but of a mixture of molecules of hydrogen chlorid and ammonia. When allowed to cool, these molecules reunite and form molecules of the original compound. This phenomenon is called *dissociation*.

EXPERIMENT 24

Dissociation

Procure a glass tube about 25 cm. long and 2 cm. bore, and close the ends by means of two closely fitting corks,



through both of which passes the stem of a clay tobacco pipe. Place in the center of the tube a piece of crystallized ammonium chlorid, and at the ends, next to the corks, place pieces of moist blue litmus paper, flat against the glass. Connect one end of the pipe with a pair of bellows. Heat the ammonium chlorid with the flame of a Bunsen burner, at the same time gently forcing air through the pipe stem. The two gases into

which the ammonium chlorid has been dissociated will pass through the porous pipe stem in different quantities; and,

by holding a piece of moist red litmus paper at the opening of the pipe, the presence of ammonia gas will be shown, while the paper inside the tube will show the presence of free hydrochloric acid gas.

Dissociation in Solutions. Dissociation takes place not only in gaseous compounds but also in solutions. In these cases it is probable, and it is believed, that the molecules are continually breaking up into atoms or groups of atoms, and then reuniting again. These atoms or groups of atoms are called *ions*. In the case of a solution of hydrochloric acid, the ions would be H and Cl, and in the case of a solution of copper sulfate they would be Cu and SO_4 . The ions are believed to be present in greater numbers in dilute solutions than in strong ones. The theory is that the difference between ions and simple atoms is, that the ions are charged electrically; thus, in a solution of copper sulfate, the Cu ions are charged positively, and the SO_4 ions are charged negatively.

EXPERIMENT 25

Dissociation in Liquids

Prepare four strong aqueous solutions (10 cc. each) as follows: — one of copper nitrate, one of copper chlorid, one of sodium chlorid, and one of sodium nitrate. In a test tube, mix the copper nitrate and the sodium chlorid solutions. In another, mix the sodium nitrate and copper chlorid. Note that the two mixtures are the same. Evidently each mixture contains copper nitrate, copper chlorid, sodium nitrate, and sodium chlorid. Besides these compounds there must be present the ions of Cu, Na, NO_3 and Cl. To destroy the equilibrium between these, add to one of the mixtures a little

powdered sodium chlorid and shake. The additional green color shows the presence of more copper chlorid, the formation of which has of course necessitated a rearrangement of the compounds and ions.

Organic Chemistry. It was formerly thought that those compounds formed by the chemical elements in living bodies were not bound by the same laws as those of the inorganic world. The principle of life was supposed in some mysterious way to govern them, and it was thought that they could not be prepared artificially. On this account, that branch of chemistry which deals with such compounds was named *organic chemistry*. When Wöhler, a German chemist, succeeded in making artificially the organic compound urea, this theory was overthrown. After many other compounds were thus made, it became no longer tenable, and we now class all chemical compounds as dependent upon the same laws. It happens that the element carbon is the most frequently present in the so-called organic compounds; so it were better to name this branch of the science *the chemistry of the compounds of carbon*. However, the name organic chemistry has clung to it, and probably always will.

EXPERIMENT 26

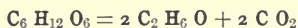
Burning of Organic Matter; Dry Distillation

a. Burn a number of organic substances such as wood, alcohol, a candle, kerosene, etc., under a cold bell-jar, and notice the formation of water. On the end of a glass rod hold a drop of lime-water in the jar, and prove the presence of carbon dioxid.

b. To a hard-glass tube about 8 mm. in diameter containing a piece of wood, attach by means of a one-holed cork

a glass exit tube. Heat the wood, and ignite the gas that escapes. What remains in the tube? Is the combustion of the wood partial or complete?

Alcohol. Sugar is a compound made up of carbon, hydrogen, and oxygen ($C_6H_{12}O_6$). Whenever a juice containing sugar is left in the open air, it decomposes, giving off carbon dioxid gas, and forming a new compound called alcohol C_2H_6O .

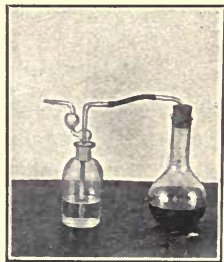


This action, called fermentation, is caused by a small organized body (in this case vegetable) called a ferment.

EXPERIMENT 27

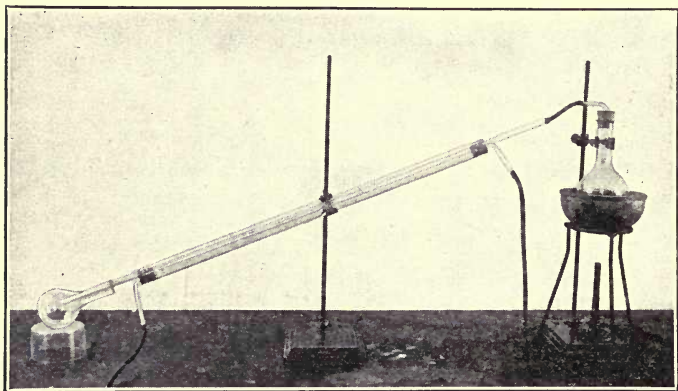
Alcohol

In a 500 cc. flask, dissolve 40 grms. of grape sugar in 250 cc. of water. Add to this a little brewer's yeast, after connecting the flask with a wash bottle containing lime water. Notice the evolution of carbon dioxid, as proved by the milky color of the lime water. After the apparatus has stood long enough for the action to cease, remove the flask and place it in a water bath. Connect it with condenser, and allow the alcohol to distill over. The condenser may be made as follows.



Fit each end of a glass tube (about 60 cm. long and 2.5 or 3 cm. bore) with a two-holed rubber stopper. Let a glass tube about 75 cm. long and 5 mm. bore extend through the tube, and through both stoppers. In the other holes of the

stoppers, fit pieces of glass tubing about 10 cm. long and bent at right angles. Fasten the apparatus to a stand, and incline it at a slight angle. Allow cold water by means of rubber tubing to enter the lower end and escape from the higher.



EXPERIMENT 28

Saponification

In a porcelain dish, boil for an hour about one-eighth of a pound of lard together with a solution of sodium hydroxid (20 grms. to 125 cc. of water). While the mixture is cooling, add a strong solution of salt. The substance that solidifies on the surface of the liquid is soap.

Remark. Soaps are the alkali salts of certain fatty organic acids such as *stearic*, $C_{18}H_{36}O_2$, and *palmitic* $C_{16}H_{32}O_2$. The calcium and magnesium salts of these acids are insoluble in water; hence, when soap is used with hard water (see Exp. 35, Part I.), these compounds appear on the surface of the water as a scum.

PART III.

HISTORY, OCCURRENCE AND
INDUSTRIAL APPLICATIONS OF THE
PRINCIPAL ELEMENTS AND COMPOUNDS

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HISTORY, OCCURRENCE AND INDUSTRIAL APPLICATIONS OF THE PRINCIPAL ELEMENTS AND COMPOUNDS

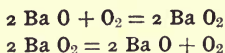
OXYGEN

History. Until 1774, the air was believed to be a simple substance. In that year, the investigations of Priestley, Rutherford, and Scheele proved that it was a mixture of two different gases. By heating mercuric oxid (see Exp. 5, Part I.), Priestley proved that this substance was composed of a gas and metallic mercury. The gas thus obtained was shown to be the same as one of the constituents of the air. In 1805, Gay Lussac proved that water was composed of two volumes of hydrogen and one of oxygen. (See Exp. 6 *c*, Part I.) The name oxygen ($\acute{o}\xi\acute{\upsilon}\varsigma$, sour, $\gamma\epsilon\upsilon\nu\acute{\alpha}\omega$, I produce) was given it by Lavoisier.

Occurrence. Oxygen is the most abundant element in nature. It constitutes 23 per cent of the atmosphere, 88.88 per cent of water, and from 44 to 48 per cent of the crust of the earth. The ores of almost all of the metals occur in the earth as oxids, or as other compounds containing oxygen.

Industrial Applications. Oxygen is used extensively for medical purposes, and in connection with hydrogen in the oxy-hydrogen blowpipe. It is usually prepared in large quantities by heating potassium chlorate (KClO_3), or by obtaining it from the atmosphere. In the latter method,

advantage is taken of the fact that at a dull red heat barium oxid (BaO) takes on oxygen, becoming barium peroxid (BaO_2), and that at a still higher temperature the oxygen thus absorbed is given off, leaving the original oxid. Theoretically this process could be continued indefinitely. In practice, however, the barium oxid is not used indefinitely, since it becomes gradually less efficient in its action. The changes are



It has been found that this reaction can be accomplished at a constant temperature by changing the pressure.

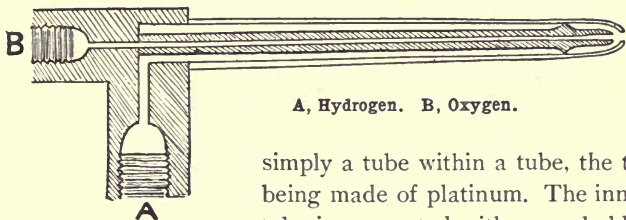
HYDROGEN

History. In the sixteenth century, Paracelsus obtained an inflammable gas by treating metals with certain acids. To this gas, Cavendish gave the name "Inflammable Air" in 1766. Later he proved that, when this gas was united with oxygen, water was formed. Lavoisier confirmed this, and gave to the gas the name hydrogen ($\nu\delta\omega\rho$, water, $\gamma\epsilon\nu\nu\acute{\alpha}\omega$, I produce).

Occurrence. Hydrogen occurs in the free state in the atmosphere of the sun, in small quantities mixed with other gases in volcanic eruptions, and sometimes in oil wells. It is also found occluded in meteoric iron and certain iron ores. In chemical combination, it is most widely distributed as water. It also occurs in combination with a number of the non-metals, and as a part of almost all organic compounds.

Industrial Applications. From its great lightness, hydrogen is valuable for filling balloons. It has often been made for military balloons by the steam and red-hot iron process. (See Exp. 15 b, Part I.)

The great amount of heat given out by the oxidation of hydrogen has made it valuable in melting refractory metals such as platinum. This is done by means of a very simple piece of apparatus called the oxy-hydrogen blowpipe. It is



A, Hydrogen. B, Oxygen.

simply a tube within a tube, the tip being made of platinum. The inner tube is connected with a gas holder containing hydrogen, while the other is connected with a similar one containing oxygen. The hydrogen is first turned on and ignited. The oxygen is then turned on until the flame burns quietly.

WATER

Water is one of the most abundant and most widely distributed compounds. From the fact that it is a solvent for so many substances, it is never found pure. When pure, it is tasteless and colorless. In large quantities, it often has a greenish or a bluish color. The great reservoir of course is the sea, from which the water that is precipitated upon the land as rain, hail, or snow, originally evaporated. Sea water contains in solution about $3\frac{1}{2}$ per cent of solid matter, most of which is sodium chlorid. The water of lakes and rivers contains various substances depending upon the locality. The most common substances are calcium carbonate and calcium sulfate. (See Exp. 31, Part I.)

Water is so widely used that it is hardly necessary to enumerate many of the ways in which it is of value. The principal uses of water, however, are for drinking,

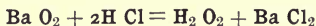
washing, and culinary purposes in the household, and for boiling and solvent purposes in the arts. Good drinking water should be as free as possible from sewage contamination and decaying organic matter. These contaminations are not only injurious in themselves, but they render the water more habitable for the germs of typhoid fever, cholera, and other malignant diseases. The disease germs are almost always carried into water by means of sewage, hence water containing the slightest trace of sewage should be especially avoided. Water that contains organic impurities is usually yellowish in color, and has a disagreeable odor.

The water supply of large cities is freed from impurities by a process of filtration through several feet of sand and broken stone extending over a large area. At least two such filters are used, and one is kept empty while the other is in use. The disease germs that are removed by one filter are destroyed by oxidation, when the water is diverted into the other.

For boiling and solvent purposes in the arts, soft water is preferable. For washing purposes, hard water may be made soft by the addition of sodium carbonate.

HYDROGEN DIOXID, H_2O_2

This compound was first discovered by Thénard in 1818. He obtained it by treating barium peroxid with dilute hydrochloric acid.



It is formed in minute traces in the atmosphere, and is sometimes produced simultaneously with the preparation of ozone. It is an oily, colorless liquid, having a bitter taste. It is usually used in dilute solutions.

Industrial Applications. Because of the fact that hydro-

gen peroxid gives up part of its oxygen readily, it is very valuable as an oxidizing agent. It has been extensively used for bleaching hair, giving to dark hair the well-known light flaxen tint. It is also used for restoring the colors to old paintings. White paint is composed largely of lead carbonate, which gradually darkens in time. This occurs on account of the action of sulfur, which forms a black lead sulfid. If hydrogen peroxid is applied to such a discolored painting, the lead sulfid is converted into lead sulfate, thus restoring in a great measure the original color. Hydrogen peroxid is also used extensively in medicine and as a disinfectant.

THE HALOGENS

The elements chlorin, bromin, iodin, and fluorin may be grouped together, since their properties are in many ways similar. This group is called the *halogen* group (*ἅλος*, salt, *γεννάω*, I produce).

CHLORIN

History. Chlorin gas was first obtained and studied by Scheele in 1774. He obtained it from hydrochloric acid and manganese dioxid. Its elementary character was proved by Davy in 1810, and he named it chlorin from *χλωρός*, meaning *greenish yellow*.

Occurrence. On account of its strong affinity for other elements, chlorin is not found in the free state. It occurs in combination chiefly with the alkali metals in sea water, and as rock salt in various localities (chiefly at Stassfurt, Germany). At Syracuse, New York, salt occurs as a brine at a depth of from 200 to 400 feet below the surface of the earth.

Industrial Applications. From its strong affinity for hydrogen, chlorin is used in the arts in the process of bleaching

cloth. In bleaching cloth, it is necessary for the cloth to be moist. The change is explained by the fact that the chlorin unites with the hydrogen of the water, leaving nascent oxygen, which in turn oxidizes the vegetable coloring matter into colorless compounds. Black colors from carbon, as for instance printer's ink, cannot be bleached. Ordinary writing ink, which is a compound of organic acids and iron, is readily decolorized.

Large quantities of chlorin are also used for disinfectant purposes.

Manufacture. It will be remembered that chlorin is made in the laboratory (see Exp. 27 *b*) by treating manganese dioxid with hydrochloric acid, thus:—



Since manganese dioxid is the costly material used, it is necessary in the manufacture of chlorin on a large scale to save the manganese chlorid formed, and in some way change it back to the oxid. In the *Weldon* process this is done. The chlorid is treated with calcium hydrate. The mixture is then heated, and a current of air is blown through. A complicated reaction that need not be entered into here ensues, and the manganese is oxidized to peroxid. Thus the original manganese can be used over again, making chlorin a cheap commercial product.

Other processes for the manufacture of chlorin are also in use.

HYDROCHLORIC ACID

History and Occurrence. Hydrochloric acid was known to Arabian alchemists in *aqua regia*. Basil Valentine in the fifteenth century wrote of an acid which he called "spirit of salt," obtained from oil of vitriol and common salt. Priestley

obtained the gas by collecting it over mercury instead of water, and called it "*marine acid air*." In 1810, Davy showed that this gas was composed of hydrogen and chlorin, and that it was not a compound of oxygen as had formerly been supposed.

Hydrochloric acid occurs in the gases which issue from some volcanoes, especially Vesuvius. It is also found in some South American rivers whose sources are in the volcanic districts of the Andes.

Industrial Applications and Manufacture. Hydrochloric acid is used in the manufacture of chlorin, ammonium chlorid, and tin chlorid, the last being extensively used by dyers.

In the manufacture of sodium sulfate, large quantities of hydrochloric acid are obtained as a by-product (see Exp. 23, Part I.). The acid fumes pass through a flue to brick chambers filled with coke or broken brick through which water is passing. The water absorbs the gas, and this solution, when collected, is the hydrochloric acid of commerce.

OXIDS AND OXY-ACIDS OF CHLORIN

There are three oxids of chlorin.

Chlorin monoxid, $\text{Cl}_2 \text{O}$

Chlorin trioxid, $\text{Cl}_2 \text{O}_3$

Chlorin peroxid, Cl O_2

They are all unstable compounds.

The oxy-acids of chlorin are

Hypochlorous acid, H Cl O

Chlorous acid, H Cl O_2

Chloric acid, H Cl O_3

Perchloric acid, H Cl O_4

These only exist in aqueous solutions. Hypochlorous and chloric acids form salts that are important commercial products.

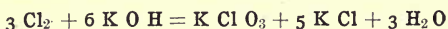
BLEACHING POWDER

If chlorin is allowed to pass into chambers containing quantities of freshly slaked lime, it is absorbed by the lime, and a compound commonly called "chlorid of lime" * is formed. This compound has strong bleaching properties because it slowly gives off chlorin. Considerable time has been given to the study of the composition of this substance, but at present no entirely satisfactory explanation has been given. It was first thought to be a salt of hypochlorous acid, but it does not show the proper percentage of chlorin. The best authority on the subject affirms that its composition is CaOCl_2 , and calls it chloro-hypochlorite. It is also called calcium oxychlorid.

"Chlorid of lime" is used in enormous quantities as a disinfectant, and for bleaching purposes. In bleaching calico and paper pulp, a two per cent solution of bleaching powder is used. The soaked product is then placed in a dilute solution of sulfuric acid, which liberates the chlorin more freely. When thoroughly bleached, the product is treated with sodium sulfite (called anti-chlor), which removes all traces of unused chlorin. This is done because otherwise the chlorin would slowly attack the fiber of the cloth or paper.

POTASSIUM CHLORATE

The potassium salt of chloric acid is a very important product. It can be made by treating warm potassium hydroxid with chlorin gas. The reaction is as follows:—

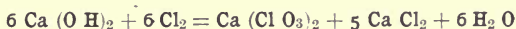


By this method, five of the six molecules of potassium

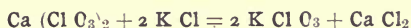
* If bleaching powder is treated with an acid, the chlorin is given off again.

hydroxid are converted into potassium chlorid; so it is better first to form the corresponding calcium salts, and then add to the solution enough potassium chlorid to change the calcium chlorate to potassium chlorate. The use of potassium-hydroxid, which is the expensive compound in the first method, is avoided in the second.

The reactions are



and



BROMIN

History. Balard discovered the element bromin in 1826. He prepared it from *bittern*, the liquid remaining after sodium chlorid has been crystallized out of concentrated sea water. He named it bromin from *βρῶμος*, a bad odor.

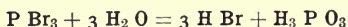
Occurrence. Bromin, as has been stated, occurs in sea water. It also occurs in the waters of many mineral springs. In these cases it is in the form of bromids of sodium, potassium, magnesium, or calcium. It also occurs in ores in combination with silver, as silver bromid.

Industrial Application. The chief use of bromin is in the manufacture of bromids and as an oxidizing agent. Its compounds are also used in photography and medicine.

HYDROBROMIC ACID

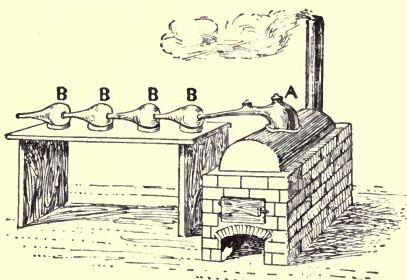
Preparation. In Exp. 28 *a*, Part I., the student made hydrobromic acid by the direct union of hydrogen and bromin. In *e* of the same experiment, it was found that the acid could not be prepared in a manner similar to that of hydrochloric acid. The best way to prepare it is to make use of the action of phosphorus bromid upon water. A

mixture of one part of amorphous phosphorus and two parts of water is placed in a flask fitted with a delivery tube and a funnel tube with a stop-cock. The flask is connected with a U tube containing a mixture of broken glass and amorphous phosphorus. Bromin is allowed to enter the flask, drop by drop, and hydrobromic acid is evolved. The phosphorus in the U tube takes up any bromin vapors that may accompany the acid gas.



IODIN

History. Courtois, in 1812, discovered iodine in the solutions of the sodium salts obtained from kelp, the ashes of seaweed. Its name is derived from *ἰοδῆς*, meaning violet-colored, on account of the violet color of its vapor.



A, Leaden still. B, Earthenware receivers.

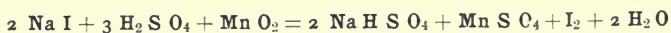
Occurrence. Iodine, like chlorine and bromine, does not occur uncombined with other elements. It is found combined with the metals in the form of iodides, both in the animal and the vegetable kingdoms. The chief source is the ashes of deep-sea weeds, which are col-

lected on the coasts of Ireland, Scotland and France. The percentage of iodine in seaweed is, however, very small.

Extraction from Seaweed. The seaweed is first dried and then carbonized. The carbonized material is next treated with water, which dissolves the sodium iodide in connection

with some chlorids, bromids, and other soluble salts. These are crystallized out as much as possible. To the remaining liquid, sulfuric acid is added, which decomposes any carbonates or sulfids that may be present, and at the same time liberates some bromin and iodine. The liquid is then placed in a leaden still (see ill. p. 168), which is connected with a series of earthenware receivers. The still is heated gently, and manganese dioxid is added from time to time. The vapors of iodine distill over, are collected, and are purified by redistillation.

The reaction is



Industrial Applications. Iodine is of great use in medicine, and in the manufacture of aniline dyes. It is also used extensively in the preparation of numerous organic compounds.

Hydriodic Acid. Hydriodic acid can be readily prepared in a manner similar to that noted under hydrobromic acid. The acid itself is unimportant, but its salts, the iodids, are well-known commercial compounds. Potassium iodid, its principal compound, is used in photography, in medicine, and as a reagent in the laboratory.

FLUORIN

History and Occurrence. On account of its intense affinity for other elements, fluorine was not isolated until 1886, when Moissan succeeded in obtaining it and in studying its properties. Fluorine occurs in the mineral cryolite ($\text{AlF}_3 \cdot 3\text{NaF}$), and as calcium fluorid (CaF) in fluor-spar. It also occurs in small quantities in the bones and teeth of animals.

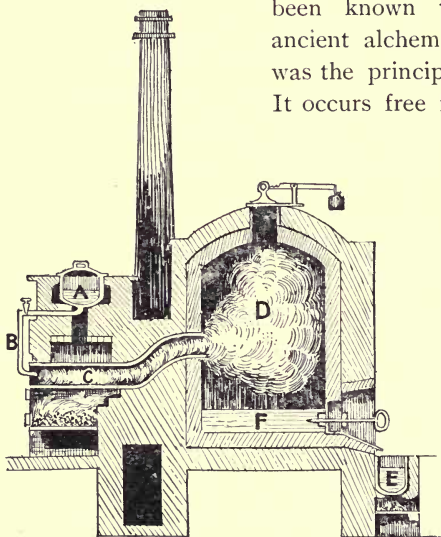
Hydrofluoric Acid. Hydrofluoric acid is made on a large scale in the same manner as in Exp. 31 b, Part I., except

that a leaden retort is used, containing the mixed calcium fluorid and sulfuric acid. The fumes that are evolved are collected in a tube immersed in a freezing mixture. This form of the acid can be stored in dilute form in gutta-percha or paraffin bottles; but if the acid is anhydrous, it can be stored only in platinum bottles, and then only at a temperature below 15°C .

SULFUR

History and Occurrence. From earliest times sulfur has been known to mankind. The ancient alchemists believed that it was the principle of combustibility. It occurs free in the neighborhood

of volcanoes, especially in Sicily, where most of the sulfur of commerce is obtained. In combination with other elements, it is found in the gases emanating from volcanoes. It is also found in the sulfates* and sulfids,† and in organic compounds.



A, Iron pot containing molten sulfur. B, Pipe leading to cylinder C. D, Brick chamber. F, Molten sulfur. E, Receiver.

* The most important are gypsum (calcium sulfate, $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$), and heavy spar (barium sulfate, BaSO_4).

† The most important are iron pyrites (iron persulfid, FeS_2), galenite lead sulfid (PbS), sphalerite (ZnS), and cinnabar (HgS).

Extraction. The free sulfur, as obtained in the native state, contains earthy impurities, stones, etc. These are removed at the mines by piling the sulfur in heaps, and setting fire to it at the bottom. The heat of the burning sulfur melts the rest of it, which runs to the bottom and is collected. This is shipped away to other countries, where it is further purified. The last purification is done by melting the sulfur in an iron pot from which it is allowed to run into a cast iron retort. The retort is heated, and the sulfur vapors are led into a brick chamber. The first sulfur that enters the chamber condenses to the solid form as a fine powder known in commerce as *flowers of sulfur*. When the walls of the chamber become hot, the fumes condense to a liquid, which is drawn out into wooden molds, and when cool, forms the *roll brimstone* of commerce. (See ill. p. 170.)

HYDROGEN SULFID

History and Occurrence. In 1777, Scheele obtained this gas by heating sulfur in hydrogen gas. It is found in nature in volcanic gases, in certain springs, and wherever organic matter decomposes.

Uses. Hydrogen sulfid is used mainly in the chemical laboratory as a reducing agent and to precipitate certain metals as insoluble sulfids.* It is also extensively used medicinally wherever springs are found whose waters contain it in solution.

The presence of hydrogen sulfid in eggs, in vulcanized rubber, and in many other organic substances containing

* The presence of hydrogen sulfid in water or in gases can be determined by means of the precipitates formed with solutions of the soluble salts of various metals.

sulfur, may be shown by its blackening a silver coin. This is due to the breaking up of the hydrogen sulfid, which, decomposing very easily, yields nascent sulfur. The black spot on the silver is silver sulfid.

SULFUR DIOXID

History and Occurrence. In the writings of Homer, mention is made of using burning sulfur for fumigation purposes; and Pliny tells of its use for purifying cloth. Priestley, however, first prepared pure sulfur dioxide, in 1775. Like hydrogen sulfid, sulfur dioxide occurs in nature in volcanic gases.

Manufacture and Industrial Applications. Sulfur dioxide can be economically manufactured in three ways, — by heating sulfuric acid and charcoal, by burning sulfur in a furnace, and by roasting iron pyrites (FeS_2). In the manufacture of sulfuric acid, where sulfur dioxide finds its greatest use, the last two methods prevail. Sulfur dioxide is valuable also on account of its bleaching properties. Substances that would be injured by chlorine, such as silk, wool, straw, etc., are bleached by this gas. The material is moistened and hung or placed in chambers into which the fumes of burning sulfur are allowed to enter. Colors acted upon by chlorine cannot be restored, but colors acted upon by sulfur dioxide can be restored by treating them with dilute sulfuric acid or with alkalis. Sulfur dioxide is used also to remove chlorine from freshly bleached goods. The sulfur dioxide and the chlorine here form sulfuric and hydrochloric acids, which are afterwards removed by washing. Usually sodium sulfite, the sodium salt of sulfurous acid, is used instead of sulfur dioxide. Sulfur is often burned in sick rooms, since the re-

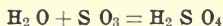
sulfating sulfur dioxid acts as a disinfectant. Some authorities claim, however, that as a disinfectant its value is very limited.

SULFURIC ACID

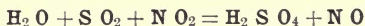
History and Occurrence. Basil Valentine first described sulfuric acid in the 15th century, though it was probably known before that time. The acid was first made from ferrous sulfate ($\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$), which, when heated, first gives off water and then sulfur trioxid (See Exp. 9 e, Part I.). These in turn unite, forming an acid called fuming sulfuric acid. In 1770, Roebuck, of Birmingham, proposed the lead-chamber process, which is still used.

Sulfuric acid is found in the free state in some rivers and mineral springs.

Manufacture. In Exp. 9 f, Part I., it was found that sulfuric acid could be made by adding water to sulfur trioxid.

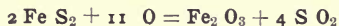


On a large scale, however, this method is impracticable, and a modification of the method, described in Exp. 39 b, Part I., is used. In Exp. 39 b, Part I., the reaction that takes place is

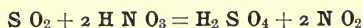


The NO_2 is formed by the oxidation of the NO obtained from the flask. Thus NO acts only as a carrier of oxygen.

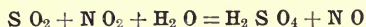
Sulfuric acid is *manufactured* in a series of lead-lined chambers. The lead is used because it is the only inexpensive metal that sulfuric acid, below a certain strength, does not attack. In this process (called the *lead-chamber process*), the sulfur dioxid is obtained by burning sulfur or by roasting iron pyrites in a furnace. The latter method is usually used.



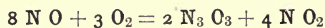
The mixture of sulfur dioxid with air, both obtained from the furnace, passes through a flue into a tower containing loosely packed coke. This is called a *Glover's tower*. From the Glover's tower the mixed gases pass into a lead-lined chamber into which enters a jet of steam; and thence they pass to a second similar chamber into which nitric acid flows over a number of corrugated porcelain cones. These are used to spread the nitric acid over as much surface as possible. Here a complicated reaction ensues, which, without going far from the truth, may be stated briefly



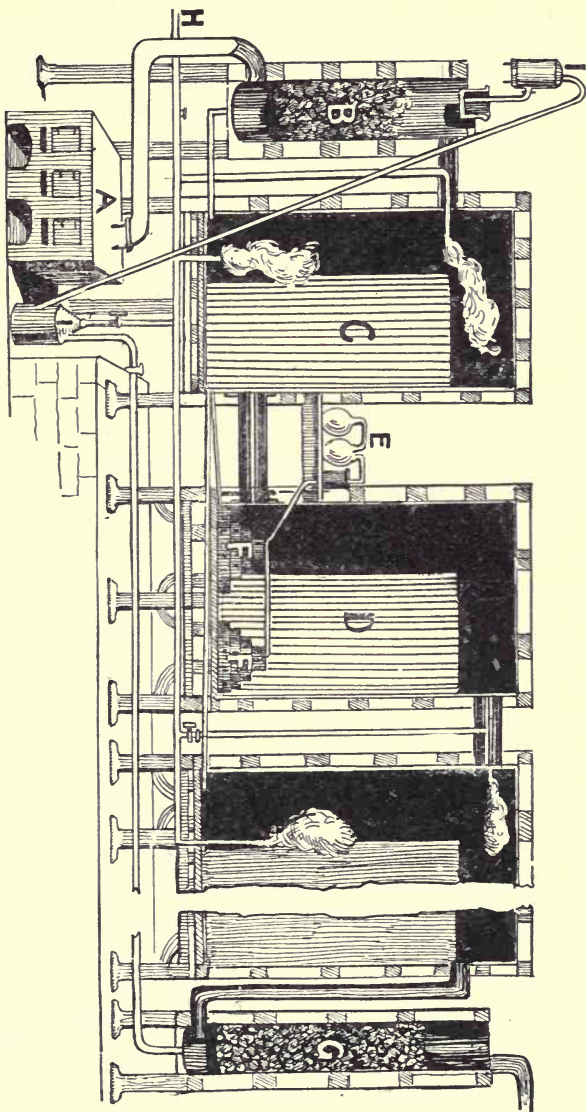
The sulfuric acid thus formed contains unused nitric acid and oxids of nitrogen. The liquid is allowed to flow back into the first chamber, where it meets a fresh supply of sulfur dioxid, steam, and air. The gaseous products (SO_2 , NO_2 , and steam, not yet combined in the second chamber) then pass into other lead chambers into which steam jets enter. Here the following reaction takes place:—



The NO is also oxidized to NO_2 and N_2O_3 ,



The last of these chambers is connected with a tower, called a *Gay Lussac tower*, filled with loosely packed coke through which slowly trickles a stream of strong sulfuric acid. This acid dissolves the unused oxids of nitrogen that would otherwise escape with the waste nitrogen left from the air, and is then collected at the bottom and forced to the top of the Glover's tower by means of compressed air. It there trickles down through the coke, and meets a fresh supply of sulfur dioxid and air, thus making more sulfuric acid.



A, Furnace for roasting sulfur or pyrites. B, Glover's tower. C and D, Lead chambers. E, Nitric acid reservoirs. F, Corrugated porcelain cones. G, Gay Lussac tower. H, Steam pipe. I, Apparatus for forcing acid to J.

The sulfuric acid thus obtained is about 65% pure, and has a specific gravity of about 1.5. It is removed to leaden pans, and evaporated until it reaches a specific gravity of 1.75. After it reaches this strength, lead cannot be used, because it is attacked by the acid. Further concentration is reached by evaporating it in platinum dishes until the acid reaches a specific gravity of 1.84. Instead of platinum, iron vessels have recently been used, since iron is not attacked by the concentrated acid.

Industrial Applications. Sulfuric acid is the most widely used of all chemicals. Thousands of tons are produced annually. It is a necessity in the manufacture of most acids and salts. The manufacturers of soda ash, glass, soap, and fertilizers, are great consumers of sulfuric acid, in fact, there is scarcely a manufacturing plant of any kind that does not use the acid for some purpose or other.

SELENIUM AND TELLURIUM

History and Occurrence. There are two elements, selenium and tellurium, that are very similar to sulfur both in their chemical properties and in their compounds. Both are rare elements. Selenium generally occurs combined with copper, lead, or silver. It is a dark, reddish-brown, crystalline solid in one of its allotropic forms, while in the other it is a dark, gray solid. Tellurium usually occurs as a tellurid of silver, gold, or bismuth, and is a white, brittle solid of metallic luster. The former element was discovered by Berzelius in 1817, the latter by Klaproth in 1798.

NITROGEN

History. In 1772, Rutherford showed that if animals breathe in a confined volume of air, and the carbon dioxid

thus formed is removed by means of lime water, a gas is left that does not support combustion. In the same year, Priestley showed that the same thing is true of confined air in which carbon has been burnt. Lavoisier gave it the name of *azote*, but the name nitrogen, suggested by Chaptal, was generally accepted.

Occurrence. About four-fifths of the air is nitrogen (see Exp. 36, Part I.). This gas is also a constituent of a large number of organic compounds, and occurs in combination with hydrogen as ammonia, and with metals and oxygen as nitrates.

AMMONIA

History. The early alchemists were familiar with the odor of the salt now called ammonium carbonate. Basil Valentine showed that, when an alkali acts upon sal-ammoniac, a strong smelling solution is obtained. The name *sal-ammoniac* was given to a salt that, in the seventh century, was brought to Europe from Asia. It was later obtained by the dry distillation of animal matter, such as hoof, hair, horn, etc. The carbonate of ammonium obtained in this way was neutralized with hydrochloric acid, and ammonia was obtained from the resulting ammonium chlorid. In 1774, Priestley heated sal-ammoniac with lime, and collected the resulting gas over mercury. By passing electric sparks through the gas, he found it to be a compound of hydrogen and nitrogen.

Occurrence. Ammonia occurs in small quantities in the air, where it is believed to be formed by electric discharges. It is formed also by the decay of animal matter. Ammonium salts occur in the soil and in certain mineral waters.

Manufacture and Applications. Practically all the ammonia of commerce is obtained as a by-product in the manufacture

of coal gas. When coal is heated in the absence of air, the nitrogen passes off in combination with hydrogen as ammonia. This dissolves in the condensed moisture which is formed at the same time. From this liquid, ammonia is obtained.

Ammonia has come to be one of the most important of commercial chemicals. This has been due in a great measure to its use in making artificial ice. Ammonia gas condenses to a liquid under a pressure of seven atmospheres at the ordinary temperature. When it is allowed to evaporate, a great amount of heat is rendered latent, hence its value in refrigeration. It is also used in the preparation of a great number of compounds, some of which will be mentioned later. The value of compost as a fertilizer is due to the ammonia evolved by the decomposition of the animal matter in it.

OXIDS OF NITROGEN

Nitrogen forms five oxids as follows:—

Nitrous oxid, $\text{N}_2 \text{O}$ (See Exp. 46, Part I.)

Nitric oxid, N O (See Exp. 39, Part I.)

Nitrogen trioxid, $\text{N}_2 \text{O}_3$

Nitrogen peroxid, N O_2

Nitrogen pentoxid, $\text{N}_2 \text{O}_5$

Only two are important, the nitrous oxid and the nitric oxid. The former is the well-known laughing gas, used by the dentist; while the latter finds its chief use in the manufacture of sulfuric acid.

ACIDS OF NITROGEN.

Of the three acids of nitrogen, viz.,

Hyponitrous acid, H N O

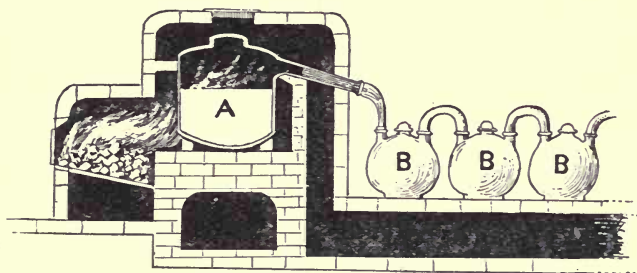
Nitrous acid, H N O_2

Nitric acid, H N O_3

only nitric acid is important. Some salts of the other two, however, are extensively used.

NITRIC ACID

History. Geber, the Arabian alchemist, seems to have been the first person to make nitric acid. It was called *aqua fortis* by the alchemists. The present method of making nitric acid from a nitrate and sulfuric acid was first spoken of in the writings of Glauber in the 17th century. Cavendish, in 1785, determined definitely that it was a compound of hydrogen, nitrogen, and oxygen.



A, Retort. B, Earthenware receivers.

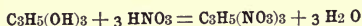
Occurrence. Nitric acid is probably found in the atmosphere, since we might expect it to be formed there by electric discharges. It occurs plentifully on the earth in the form of nitrates, the chief of which are potassium nitrate (saltpeter) and sodium nitrate (Chili saltpeter). Whenever nitrogenous organic matter decomposes in the air, the nitrogen takes the form of ammonia. If an alkali is present, a slow oxidation takes place, and forms a nitrate of the alkali present. In this manner, the nitrate beds of India, Chili, and Peru have been formed.

Manufacture. Nitric acid is made for commercial use in the same way as in the laboratory (see Exp. 37, Part I.). Large cast-iron retorts, protected inside by a lining of clay, are used in which sodium nitrate and sulfuric acids are placed. The nitric acid formed distills over into earthenware bottles. (See ill. p. 179.) It is then purified by redistillation with an equal volume of sulfuric acid. It is then rendered colorless by means of a current of dry air made to pass over it while it is gently warming. In some establishments, instead of the earthenware bottles, a system of tubes called "Hart's condensation tubes" are used, in which the acid is condensed and purified.

Industrial Applications. As has been stated before, nitric acid is extensively used in the manufacture of sulfuric acid, and, in that process, is the source of the necessary nitric oxid. Another application of nitric acid is in the manufacture of high explosives such as nitroglycerin,* gun cotton,† etc. It is used also in the preparation of metallic nitrates, and in the manufacture of many dyes. Its salt, potassium nitrate, is used in the manufacture of gunpowder.

* When glycerin, $C_3H_5(OH)_3$, is treated with a mixture of concentrated sulfuric and nitric acids at the proper temperature, a yellow oil is obtained. This oil is called *nitroglycerin*. It is very unstable, and is especially sensitive to concussion. It may be burned in the open air but explodes if heated too highly. *Dynamite* is a certain peculiar kind of earth saturated with nitroglycerin, and can be handled with comparative safety by persons understanding its use.

The action of the acid on the glycerin may be represented by



The sulfuric acid combines with the water formed in the operation.

† *Gun cotton* is made by treating cotton fiber with a mixture of concentrated sulfuric and nitric acids. The reaction is similar to that of nitroglycerin.

ARGON

History. Argon was discovered in the air in 1894 by Rayleigh and Ramsey. It was found that nitrogen obtained from the air, and nitrogen obtained from chemicals did not weigh the same. On account of this fact, the suspicion arose that there might be some hitherto unknown element or elements present in the air. After long experimentation, the water, carbon dioxid, oxygen, and nitrogen were removed from air, and a new, inert, colorless gas, having a density of 20, was obtained. To obtain the new gas, the moisture and carbon dioxid were removed from the air, after which the oxygen was removed by passing the dry air over red-hot copper. The nitrogen was then removed by passing the remainder through a tube containing magnesium turnings heated white-hot.

It has since been found by the same experimenters that the gas obtained by the above process is not a single gas, but that there is present probably a number of different gases, very similar in their properties. Of these, *neon*, *metargon*, and *xenon* have been separated.

HELIUM

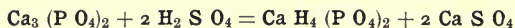
While investigating argon, Rayleigh discovered in certain rare minerals an element that had hitherto been known only to exist in the sun. It had already been called *helium*. Its atomic weight was found to be about 4.

PHOSPHORUS

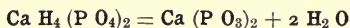
History and Occurrence. Brand claimed to have discovered phosphorus in 1669. From its property of becoming luminous in the dark, its name is taken from the words $\phi\omega\varsigma$, light,

and $\phi\acute{\epsilon}\rho\omega$, I bear. The element never occurs free in nature. It is widely distributed, however, in combination. Many rocks contain phosphorus compounds. When these rocks disintegrate, the soils that are formed become very productive on account of the presence of phosphorus. Vegetation, in turn, is transformed into animal matter. Here the phosphorus appears mostly in the bones as calcium phosphate. A fossil substance called "caprolites" also contains calcium phosphate, and it is from this and from bones that the phosphorus of commerce is derived.

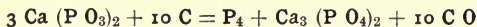
Manufacture. In the manufacture of phosphorus from bones, the first step is to remove the non-phosphorus-bearing organic matter. This is done by burning. The ash is then treated with sulfuric acid, which changes the insoluble calcium phosphate to soluble acid calcium phosphate, according to the following reaction :



After the calcium sulfate has been removed, the acid calcium phosphate solution is evaporated to a syrup and mixed into a paste with charcoal powder. This is heated to redness in an earthenware retort, whose mouth dips under water. Water is at first driven off thus : —



The metaphosphate thus formed then reacts with the charcoal as follows : —



The phosphorus distills over into the water, is collected, and is purified by redistillation. Sometimes sand is added together with charcoal, in which case all the phosphorus is obtained.

The phosphorus obtained by the above method is the yellow variety. The red amorphous phosphorus is made by

heating, for a number of days, the yellow variety at 240° C. in an iron vessel having only a small opening. Not all the yellow phosphorus changes over. The unchanged phosphorus is dissolved out of the cooled mixture by carbon bisulfid, or is removed by means of sodium hydroxid which acts upon the yellow phosphorus and forms phosphoretted hydrogen.

Industrial Applications. Phosphorus is used in the arts principally in the manufacture of matches. The old friction matches were made by first dipping the wood into melted sulfur, cooling it, and then tipping it with a mixture of phosphorus, glue, and some oxidizing agent. The so-called safety matches do not contain phosphorus. Instead of this they are tipped with antimony sulfid. The box is painted with a preparation of red amorphous phosphorus, antimony sulfid, and glue. Sometimes manganese dioxid or some other oxidizing agent is used in this mixture.

Phosphin. There are three compounds of phosphorus and hydrogen, only one of which, phosphin (PH_3), will be noted here. If phosphorus is heated in a solution of sodium or potassium hydroxid, a gas is liberated that ignites spontaneously. If the gas is led under water, the bubbles on reaching the surface burst into flame and form white rings of phosphorus pentoxid. The gas is colorless, has a very disagreeable odor, and is poisonous. It forms a class of substances, called phosphonium compounds, analogous to those formed by ammonium.

Phosphoric Acid. The phosphoric acid of commerce is made from bones. The main value of phosphoric acid lies in its salts, which are the principal constituents of fertilizers. For the purpose of making fertilizer, bone ash is treated with sulfuric acid to obtain the soluble acid calcium phosphate, $\text{CaH}_4(\text{PO}_4)_2$.

ARSENIC

History and Occurrence. Arsenic compounds were known to the alchemists, who thought they could use them in transmuting the metals. Brandt, in 1773, first showed that white arsenic was the calx of a metal. Later, its position as an oxid was established.

Arsenic occurs in the free state in nature, but usually it is found in combination with iron, cobalt, nickel, or sulfur.

Applications. Arsenic is used in the manufacture of certain pigments. In the manufacture of glass, arsenic oxid is used to remove the green tint given by ferrous hydroxid, which by its action becomes ferric oxid. In the manufacture of shot, arsenic has the effect of hardening the lead. In medicine, it is sometimes used as a tonic. For preserving skins, it is valuable to the taxidermist. In the form of Paris green (CuHAsO_3), it is valuable as a poison for the destruction of insects.

ANTIMONY

History. Mention is made in the Scriptures of the metal we call antimony. In early times the sulfid was used by the women of the East to paint their eyebrows. Pliny called it *stibium*, although in Latin it was also known as *antimonium*.

Occurrence and Applications. While antimony occurs sometimes free in nature, it is usually found as the sulfid (Stibnite, Sb_2S_3). When the ore is roasted in the presence of air, the oxid (Sb_2O_3) is formed. This, mixed with carbon and ignited strongly, gives the metal.

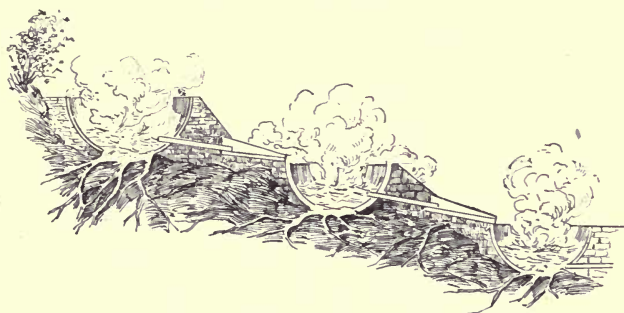
Antimony is used chiefly in making alloys. It hardens the alloy, and also, by its property of expanding when cooled, makes it invaluable in the manufacture of type metal. Type

metal is composed of about two parts lead, and one part each of antimony and tin. Antimony is used also in making white metal, pewter, and Britannia metal. Finely divided antimony, prepared by zinc from the chlorid, is called *antimony black*, and is used for giving a metallic appearance to plaster casts, statues, etc. In medicine its compound, potassium antimony tartarate (tartar emetic), is used as an emetic.

BORON

History and Occurrence. Gay Lussac, Thénard, and Sir Humphry Davy obtained boron in the elementary state in 1808. There are two varieties, the amorphous and the crystalline boron.

It never occurs free, but is found as boric acid (H_3BO_3) in Tuscany, and as sodium salts in California and Thibet.



Boric Acid, H_3BO_3 . In Tuscany, jets of steam containing boric acid escape to the surface of the earth from subterranean sources. Brick basins are built around these steam jets. The heat causes the solution obtained to evaporate. When sufficiently strong, the solution is removed and allowed to crystallize.

Uses. Boric acid is a valuable antiseptic, and for this

reason is used extensively in surgery. It is also sometimes used in preserving perishable foods.

Borax, $\text{Na}_2\text{B}_4\text{O}_7$ is a valuable compound, and is much used as a flux by tin-smiths and copper-workers. It gives a clean, bright, metallic surface on account of its property of dissolved metallic oxids.

CARBON

History. Carbon has of course been known from ancient times, but its allotropic forms were not understood until the end of the 18th century, and its relation to organic chemistry not until somewhat later. In the form of graphite, it was known to the alchemists. Graphite pencils were first made in 1565. In 1772, Lavoisier showed that diamond and charcoal were chemically identical. He burned a diamond and obtained carbon dioxid. Tennant, in 1776, proved that like weights of charcoal, graphite, and diamond give like weights of carbon dioxid.

Occurrence. Carbon occurs as diamond principally in India, South Africa, and Brazil. It is usually transparent or slightly tinged with yellow, although it is sometimes found red, green, blue, or even black. As graphite, carbon occurs widely distributed, being found mostly in England, Siberia, Ceylon, Canada, New York, and California.

Graphite is used in the manufacture of lead pencils, crucibles, and various lubricants. It is also used in foundries for facings, and in electrotyping.

Amorphous Carbon. *Coal.* As *mineral coal*, carbon is found in almost all countries of the world. The most, however, is found in England and the United States. Coal is all that is left of the great primeval forest that covered the earth long before the advent of man. When vegetable

matter decays in the absence of air, in the earth, or under water, it gives off gases, and a substance similar to coal remains. In some parts of the earth, the trees of the ancient forests were buried by earthy material, and underwent the process of decay for millions of years. Upon the completeness of this process depend the different kinds of coal. Where the decomposition has been most thorough, *anthracite* coal results. *Bituminous* coal is rich in hydrocarbons, which it gives off when heated. *Cannel coal* is a variety of bituminous coal, and is especially rich in hydrocarbons. *Brown coal* and *peat* belong to a later geological period than the others mentioned.

Charcoal. Charcoal is obtained by the partial combustion of wood. The wood is arranged in a pile, and covered with earth. It is then ignited, and allowed to burn. On account of being protected from the air, the combustion is imperfect, and only the volatile part of the wood is driven off.

Animal Charcoal. By heating bones and other kinds of animal refuse in iron retorts, a variety of charcoal that contains calcium phosphate is obtained. It is largely used as a filter for removing vegetable coloring matters from liquids.

Coke and Gas Carbon. In the retorts of the gas works, after the volatile products have been driven off from the coal, there remains behind a gray, porous solid that is called *coke*. It is also made by burning coal in ovens so as to burn out only the volatile products. In the upper parts of the gas retorts, there remains a dull black mass known as *gas retort carbon*. It is used in making carbon plates for electric batteries, and formerly for making pencils for arc-light lamps. Gas retort carbon is the purest form of amorphous carbon.

Lamp black is made by burning turpentine and other oils

rich in hydrocarbons, and collecting the smoke. It is used chiefly in printer's ink and black paint.

Hydrocarbons. In Pennsylvania, in Ohio, and in the Caucasus, an oily liquid called *petroleum* is found in the earth. This liquid is composed mostly of a mixture of hydrocarbons. These are partly gaseous, partly liquid, and partly solid. The petroleum is distilled, and the resulting mixtures are afterwards washed with water and alkali. Gasolene, naphtha, benzine, kerosene, and paraffin, are among the principal substances obtained. These are all mixtures of various hydrocarbons.

Natural gas occurs in the earth in large quantities, usually near coal beds. Its formation is supposed to be due to the dry distillation of coal in the interior of the earth.

The hydrocarbons form a definite series of compounds, each hydrocarbon of the series differing from the next by an atom of carbon and two atoms of hydrogen, i.e., by CH_2 . The first hydrocarbons of the simplest groups are

Methane, C H_4 (See Exp. 33 *e*, Part I.)

Ethylene, $\text{C}_2 \text{H}_4$ (See Exp. 33 *f*, Part I.)

Acetylene, $\text{C}_2 \text{H}_2$

METHANE OR MARSH GAS

History and Occurrence. This gas is mentioned by Pliny, and Basil Valentine speaks of its presence in mines. In 1785, Berthollet proved that methane contained both hydrogen and carbon, and in 1805 Henry showed the difference between methane and ethylene.

Methane occurs in nature in mines, and wherever vegetable matter is decaying under water; hence its common name, marsh gas. Miners call it fire damp. Natural gas is rich in methane.

ETHYLENE

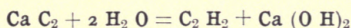
History and Occurrence. In the 17th century, Becher discovered ethylene by heating alcohol with sulfuric acid, but it was not until Henry took up the study of methane and ethylene that its composition was understood. It occurs in the gases that emanate from oil wells, and is the principal constituent of coal gas.

ACETYLENE

History and Application. Acetylene was discovered by Edmund Davy in 1836. It is formed by the incomplete combustion of other hydrocarbons, such as ethylene, coal gas, etc. For instance, when the flame "backs down" in a Bunsen burner, acetylene is formed.

Acetylene is rich in carbon, burns with a bright, white flame, and, if it can be produced cheaply enough and regulated with perfect safety, is likely to become one of the chief sources of illumination.

It is now made by the action of calcium carbide (CaC_2) upon water.

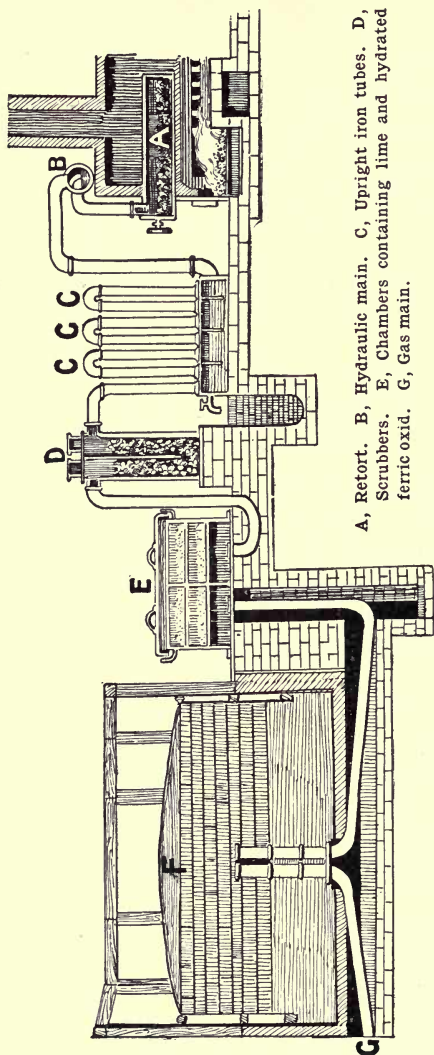


ILLUMINATING GAS

William Murdock, a Scotchman, first saw the practicability of making illuminating gas from coal. This was in 1792, but London was not lighted by gas until 1812, nor Paris until 1815.

When coal is heated in the absence of air, three classes of products are formed,

1. Illuminating gas.
2. Coal tar, a thick, oily, strong-smelling liquid.



A, Retort. B, Hydraulic main. C, Upright iron tubes. D, Scrubbers. E, Chambers containing lime and hydrated ferric oxid. G, Gas main.

3. Ammoniacal liquor, containing ammonia and various other compounds in solution.

When the coal is heated in the gas retorts, the three classes of products above mentioned pass through a pipe that dips below water in another large pipe called the *hydraulic main*. The water absorbs some of the soluble gases, and takes up some of the ammoniacal liquor and tarry products. From the *hydraulic main*, the gas passes through a series of upright iron pipes beneath which is water. As the gas passes through these pipes, it cools, and another portion of the tarry products and ammoniacal liquor is deposited, which runs down into

the water. From this pipe, the gas passes to a series of towers, called *scrubbers*, filled with coke. It passes into the bottom of one tower, through the coke, and then at the top meets a spray of ammonia water. In like manner, it passes through the next tower, and so on. The ammonia water removes most of the hydrogen sulfid. There is still left in the gas some hydrogen sulfid, carbon bisulfid, and carbon dioxid. To remove these, the gas passes through a series of chambers containing lime or hydrated ferric oxid. The gas is then ready for use, and passes into the gas holder.

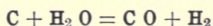
CARBON MONOXID

History and Occurrence. This compound (see Exp. 33, Part I.) was studied by various investigators during the last part of the 18th century, but it was some time before its true character was explained. It does not occur free in nature, but it is formed wherever carbon burns in an insufficient supply of air.

Applications. The principal applications of carbon monoxid are in the manufacture of water gas and in the reduction of metals from their ores.

WATER GAS

In the manufacture of water gas superheated steam is passed over red hot anthracite coal, giving the gases carbon monoxid and hydrogen.



Both of these gases burn with a blue flame, hence they must be enriched with some gas that burns brightly.

CARBON DIOXID

History and Occurrence. Van Helmont, one of the early investigators, held, as early as the beginning of the 17th century, that this gas was different from air. He called it "gas sylvestre." Its chemical nature was first determined by Lavoisier.

Carbon dioxid occurs free in the air and in combination with metals as carbonates. It is also found in some mineral waters.

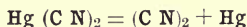
Manufacture and Applications. For commercial purposes, carbon dioxid is made by treating sodium carbonate or calcium carbonate with acids. Under 38.5 atmospheres pressure and at 0° C., this gas condenses to a liquid. It is then stored in steel cylinders, from which it can be taken at will. Its principal use is in the manufacture of aërated beverages.

CARBON BISULFID

This compound (CS_2) does not occur in nature, but is made by leading vapors of sulfur over red-hot charcoal. It is extensively used in the arts as a solvent for rubber, phosphorus, sulfur, iodine, and many oils and gums.

CYANOGEN

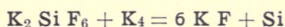
By heating mercuric cyanid, a colorless gas is obtained that burns with a beautiful purple flame, and has the odor of peach kernels.



It combines with hydrogen, forming hydrocyanic acid (HCN), commonly called *prussic acid*. This is such a deadly poison that it is used only in dilute solutions.

SILICON

History and Preparation. Silica, or sand, was believed to be a compound body long before Berzelius first obtained impure silicon by fusing together iron, carbon, and silica. It is best prepared by heating potassium silico-fluorid and metallic potassium in an iron tube.



Silicon is a brown powder that, when heated in the air, burns to silicon dioxid. If this amorphous silicon is fused with zinc, it forms dark glittering crystals of silicon that may be obtained by dissolving away the zinc with an acid.

Occurrence. Next to oxygen, silicon is the most abundant element in nature. It occurs in combination with oxygen in quartz, and in the form of silicates of the metals.

SILICON DIOXID

Quartz is the purest form of silicon dioxid. Sand and sandstone are other forms. Silicon dioxid is used in the arts in the manufacture of glass, porcelain, and in pottery.

GLASS

The Egyptians are believed to be the first people that manufactured glass. On Egyptian tombs are found pictures of glass blowers carrying on their vocation. During the Middle Ages, Venice was famous for its glass manufactures; but after their decay the art passed to the workmen of Bohemia. The art of glass making has steadily improved, until at present some of our most beautiful and marvelous works of art are due to the glass-worker's skill. Glass is a

mixture of various silicates, especially of the alkalis and alkaline earths.

The various kinds of glass are divided into classes according to their ingredients.

1. Bohemian glass.
2. Crown or window glass.
3. Common green or bottle glass.
4. Flint glass or crystal.

Bohemian glass is a silicate of potash and lime, is fusible with difficulty, and withstands chemical reagents better than any other kind.

Crown glass is a silicate of soda and lime. It is more readily fusible, and is more easily acted upon by chemicals than is Bohemian glass.

Bottle glass is a silicate of soda and lime, mixed with the oxids of aluminum and iron. The green color is due to the iron. This color varies from green to brown, depending upon impurities.

Flint glass is potash-lead silicate. It is the softest kind of glass, and has a bright luster and high refractive power. On this account, it is used in making lenses for optical instruments.

Glass can be colored by means of various metallic oxids. Gold compounds give it a beautiful ruby tint. Cuprous oxid colors glass an intense red, while cupric oxid colors it green. Cobalt gives blue, manganese gives violet. Black is obtained by the addition of sesqui-oxid of iridium.

PORCELAIN

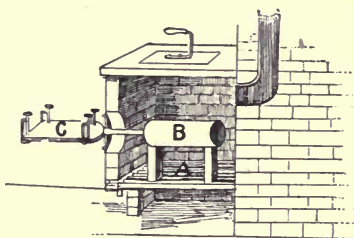
Porcelain differs from glass in that it is made of kaolin, a silicate of aluminum, $\text{Al}_2(\text{SiO}_3)_3$. Porcelain is glazed after it has been "fired" for the first time. By one method,

the ware is dipped into a glazing material, usually finely powdered quartz and feldspar, is then dried, placed in earthenware pots, and heated in a furnace. The glaze fuses and spreads over the surface of the ware. The ware is then allowed to cool in the furnace to anneal it; otherwise it would be brittle on account of the unequal tension of its different parts. Ordinary pottery is simply baked, and not glazed. Bricks are simply baked clay. The red color is due to the presence of iron silicate.

POTASSIUM

History and Occurrence. Until 1807, the alkalis were believed to be simple substances. The discovery of potassium by Davy dispelled this idea, and the true character of these metals was shown. Davy isolated potassium by decomposing potash by means of a strong electric current.

Potassium is found widely distributed in nature. It forms from 2 to 3 per cent of our granite rocks. In combination with chlorine, as potassium chlorid, it is found in the earth in considerable deposits. No vegetable growth is possible without potassium; hence all fruitful soils contain it.



A, Grate. B, Retort. C, Receiver.

Manufacture. Potassium is obtained by reducing potassium carbonate with carbon. A mixture of potassium tartarate and potassium carbonate is first heated. The tartarate decomposes into carbonate and carbon. This mixture is then placed in a wrought iron mercury bottle, and is heated.

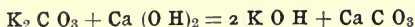
Potassium and carbon monoxid results: $K_2CO_3 + 2 C = K_2 + 3 CO$. These two substances readily form an explosive compound at high temperatures, but this is now prevented by cooling the potassium as fast as it is produced. The retort is connected, by means of a short tube, with a very shallow receiver consisting of two thin iron plates clamped together. The potassium here condenses to a liquid, and is transferred to a vessel containing naphtha, as soon as it solidifies. (See ill. p. 195.)

POTASSIUM CARBONATE

Potassium carbonate was formerly manufactured almost entirely from wood ashes. The ashes were treated with water, filtered, and the solution evaporated. Nowadays, in addition to the above source, most of the "potash" of commerce is obtained from three sources: (*a*) from beet-root, (*b*) from the sweat of sheep, and (*c*) from potassium sulfate. (*a*) The molasses from beet-root sugar is allowed to ferment, and is then evaporated. A black mass containing the potash is then obtained. (*b*) One-third of the weight of the sweat of sheep is potassium compounds. The washings of sheep wool are evaporated to dryness, and then heated in retorts. What is left is carbon and various potassium salts, which are then separated. (*c*) Potassium sulfate is obtained as a by-product in many processes. It is converted to the carbonate by a process noted later under the manufacture of sodium carbonate.

POTASSIUM HYDROXID

Potassium hydroxid is obtained from the carbonate by treating it with slaked lime.



To a hot solution of potassium carbonate, lime is added

until, after the calcium carbonate formed has settled, the addition of hydrochloric acid causes no effervescence. The liquid is then drawn off, evaporated, and finally heated to redness in silver crucibles. From the crucibles, it is run into cylindrical molds, and thus cast into sticks.

Potassium hydroxid, together with sodium hydroxid, are used in the manufacture of soap. Soap is a salt of an alkali and an organic acid.

POTASSIUM NITRATE

This substance is found in nature exuding from the soil in warm climates. (See page 179.) It is also obtained artificially on what are called "niter plantations." A mound of chalky soil is built upon a foundation of clay. This is kept moist with the nitrogenous refuse from stables and sewers. In time the organic matter oxidizes, and, uniting with the alkalis, forms nitrates. The soil is washed from time to time to obtain these salts. From the solution thus obtained, crude potassium nitrate is separated by crystallization.

Application. The chief uses of potassium nitrate are in the manufacture of nitric acid and gunpowder.

Gunpowder. Gunpowder is a mixture of potassium nitrate, charcoal, and sulfur. The percentages are approximately, potassium nitrate 75, charcoal 15, and sulfur 10. The oxygen necessary for the burning of the charcoal and sulfur is furnished by the potassium nitrate. This combustion takes place very rapidly, forming large quantities of carbon dioxid and nitrogen; hence the explosion.

OTHER POTASSIUM SALTS

Among other potassium salts that need not be described here may be mentioned potassium chlorid (KCl, see Exp.

25 *c*), potassium bromid (KBr, see Exp. 28 *c*), potassium iodid (KI, see Exp. 30 *a*), potassium chlorate (KClO_3 , see Exp. 6 *c*), potassium sulfid (K_2S), and potassium cyanid (KCN).

SODIUM

History and Occurrence. Sodium was first obtained in 1807 by Davy in the same manner as he obtained potassium. (See page 195.) It never occurs free in nature, but in combination it is very plentiful, occurring in the sea in sodium chlorid, and on the land chiefly as sodium chlorid, nitrate, carbonate, and sulfate.

Manufacture and Uses. It is prepared in a manner analogous to that of potassium. In the case of sodium, however, there is no liability to explosions, since sodium does not form a compound with carbon monoxid.

Its chief use is in the preparation of the metals manganese and aluminum. It is comparatively cheap.

SODIUM CHLORID

Occurrence and Extraction. Sodium chlorid, as has been stated before, occurs plentifully in the sea and in various salt beds.

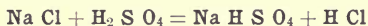
It can be obtained easily from sea water; but by far the greater amount of the salt of commerce is either mined in the solid state, as rock salt, or else is extracted from the earth in the form of brine. The brine is then evaporated.

Uses. Besides its uses for seasoning food and preserving meats, it is the basis of the great soda industry. First the sulfate is made from the chlorid, and then the other sodium compounds are obtained from the sulfate.

SODIUM SULFATE

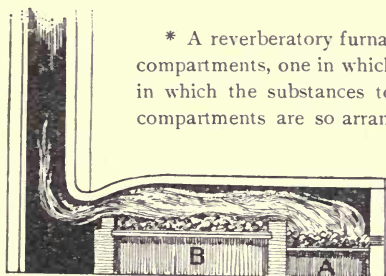
Occurrence. Sodium sulfate occurs in nature as the mineral *thénardite*. The sulfate of commerce, however, is made from common salt.

Manufacture and Uses. — Sodium chlorid and concentrated sulfuric acid are mixed in a large iron pan or retort, and this mixture is heated in a reverberatory furnace.* The mixture is first gently heated, acid sodium sulfate and hydrochloric acid being formed.



The acid gas passes through a flue into a series of towers containing coke, through which water is trickling. The water collects the acid fumes, and is drawn off at the bottom, forming the hydrochloric acid of commerce.

The mass left in the pan is then transferred to another part of the furnace, and is subjected to a higher temperature. This forms the normal sodium sulfate (see Defs. under Exp. 18), which is the "salt cake" used in the manufacture of sodium carbonate.



* A reverberatory furnace is one that has two or more compartments, one in which the fuel is burned, and others in which the substances to be treated are placed. The compartments are so arranged that the flames from the

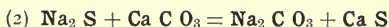
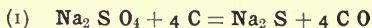
burning fuel are drawn by the draft over and deflected down upon the substances in the other compartments, thus producing the heat to act upon them. Fuel that gives long flames is generally used.

A, Firebox. B, Hearth upon which substance to be heated is placed.

SODIUM CARBONATE

There are two main processes for the manufacture of sodium carbonate. The older process is called the *Le Blanc*, while the more recent is called the *Solvay* or the *Ammonia* process.

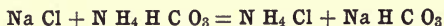
The Le Blanc Process. In the *Le Blanc* method, sodium sulfate is treated with charcoal (coal) and calcium carbonate. When heated, the sodium sulfate is reduced by the charcoal to sodium sulfid (Na_2S), which in turn acts with calcium carbonate forming sodium carbonate.



The mixture is first heated in the coolest part of a reverberatory furnace, to produce the first reaction. After a time it is placed in the hottest part of the furnace, when the second reaction takes place. The "black ash" (a mixture of sodium sulfid, Na_2S , sodium carbonate, Na_2CO_3 , coal, and lime) thus obtained is treated with water. The soluble sodium carbonate is then removed from the solution.

The Solvay Process. For many years, the *Le Blanc* was the best process known for the manufacture of sodium carbonate. It has been demonstrated, however, that the *Solvay* process is more economical.

A solution of salt is impregnated with ammonia gas in the proper proportions (a molecule of ammonia for every molecule of salt). Then into this there is led carbon dioxid until the solution is saturated. Ammonium chlorid and acid sodium carbonate are formed according to the following reaction:—



On account of its slight solubility in the liquid used, the acid sodium carbonate separates out of the solution, is collected, and dried. The liquid remaining is treated with lime, and the ammonia thus obtained is used again for impregnating a new solution of salt. The acid carbonate may be heated, and the carbon dioxide from this may also be used again, the residue left after heating being the normal carbonate.

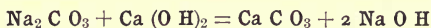
Applications. Immense quantities of sodium carbonate are used in the manufacture of glass and soap. It is also used in the preparation of other sodium compounds. The housewife uses the normal carbonate for softening water, and the acid carbonate* for cooking purposes.

SODIUM HYDROXID

After sodium carbonate has been separated from the "black ash" obtained in the *Le Blanc* process, there remains in the liquid sodium hydrate that has been formed by the action of the lime present. The solution is heated, air is blown through it, and a quantity of sodium nitrate is added. This oxidizes the sulfid present to sulfate. The solution is then evaporated to dryness, and raised almost to a red heat. The cooled product is commercial sodium hydroxid.

Sodium hydroxid is also manufactured by treating a weak solution of sodium carbonate with lime.

* Sodium bicarbonate, the acid salt, is made by passing carbon dioxide over sodium carbonate dissolved in its water of crystallization. The principal use of sodium bicarbonate is in the manufacture of baking powder. The constituents of baking powder are sodium bicarbonate and some acid, or acid salt. When mixed with dough, these constituents, by the aid of water in the dough and the heat of the oven, react upon each other. One of the products of the reaction is carbon dioxide; it is to this that the leavening process is due.



The calcium carbonate, being insoluble, is removed by filtration; and the filtrate, when evaporated, yields the hydroxid. Sodium hydroxid is purified by treating the crude product with alcohol, which dissolves the hydroxid and leaves the impurities. Such sodium hydroxid is called "soda by alcohol."

Applications. Sodium hydroxid finds its main uses in the soap factory. It is also used in purifying petroleum, carbonic acid, in the manufacture of numerous chemicals, in dye works, and in the manufacture of paper.

CALCIUM

History and Occurrence. Davy, the discoverer of sodium and potassium, first prepared calcium as a powder by electrolysis; but Matthiessen obtained the first "piece of calcium" in 1856. The metal itself does not occur free in nature, nor is it of any importance. Its various compounds, however, are very plentiful, forming, as dolomite, $\text{CaMg}(\text{CO}_3)_2$, whole mountain ranges. It is also found in large quantities as the sulfate and phosphate. The bones of animals and the shells of eggs and of mollusks are mainly composed of calcium salts. In fine, it is one of the most plentiful of the elements.

CALCIUM CARBONATE

Occurrence and Application. This compound occurs as Iceland spar, in transparent crystals. The Carrara marble is a very pure form of calcium carbonate, while ordinary limestone contains various impurities. The artificial substance is made by treating the chlorid with ammonium carbonate.

The mineral is used for building purposes, and for the manufacture of lime and cements. The beautiful marbles of Italy are used mostly by the sculptor.

CALCIUM OXID

Manufacture. The natural carbonate is heated in special furnaces, called *kilns*. Fuel is allowed to burn under the limestone, thus driving off the carbon dioxid and leaving the calcium oxid (see Exp. 21, Part I.).

Applications. The chief use of lime is in the making of mortar and cements.

Common mortar is a mixture of one part of slaked lime and three or four parts of sand made into a pasty mass. This hardens or "sets up" in a few days, but it takes years for it to harden completely. This is first due to the evaporation of the water; then, as time goes on, the lime takes on carbon dioxid from the air, and becomes calcium carbonate.

Hydraulic mortar is made from lime containing more than 10 per cent of silica, and has the ability to harden under water.

Portland cement is a hydraulic mortar made from chalk and clay. The two are ground together in water, dried, and burnt in kilns.

CALCIUM HYDROXID

Calcium hydroxid is simply slaked lime (see Exp. 20 b, Part I.).

CALCIUM SULFATE

Occurrence and Application. Calcium sulfate occurs as *gypsum* with water of crystallization, and as *anhydrite* without it. The artificial salt is formed by the action of sulfuric acid on the carbonate.

After gypsum has been heated so as to lose its water of crystallization, it is called "plaster of Paris," and possesses the property of hardening when moistened with water. It is therefore extensively used for making plaster casts, and as a cement. Gypsum is found also in some fertilizers.

CALCIUM CHLORID

Calcium chlorid is obtained as a by-product in many manufacturing processes. On account of its strong attraction for water, it is used in laboratories for the drying of gases.

BARIUM AND STRONTIUM

History and Occurrence. The elements barium and strontium were isolated by Davy in 1806, in connection with his experiments on calcium. They occur chiefly as barite (BaSO_4), witherite (BaCO_3), strontianite (SrCO_3), and celestite (SrSO_4).

Compounds. See Exps. 58 and 59, Part I.

MAGNESIUM

History and Occurrence. Magnesium also was discovered by Davy. It occurs as dolomite, $\text{MgCa}(\text{CO}_3)_2$, on the earth; as the sulfate, MgSO_4 , in mineral springs; and as chlorid, MgCl_2 , in the sea. The metal is obtained by reducing the chlorid by means of sodium.

Compounds. See Exps. 13 and 17 *b*, Part I.

Applications. Magnesium is used in many chemical operations, and as a means of artificial light in photography.

The sulfate (Epsom salts) and the carbonate are extensively used in medicine. The chlorid is used in the manufacture of cotton goods.

ZINC

History and Occurrence. The alloy brass was known to the ancients, but they did not know that it contained any metal besides copper. The discoverer of the second metal, zinc, is uncertain, although it is mentioned in writings of the 16th century. It is found plenteously as the sulfid (ZnS , zinc blende), as the oxid (ZnO , red zinc ore), as the carbonate (ZnCO_3 , calamine), and as the silicate ($(\text{ZnO})_2\text{SiO}_2$, H_2O , electric calamine).

Extraction. The ore is first roasted to convert it into oxid. The oxid is then mixed with carbon. It is then placed in cylindrical fire-clay retorts about three feet long and eight inches in diameter. These retorts are arranged in tiers and set slantingly. To the open end of each retort is joined a conical receiver about 10 inches long, extending downwards. The retorts are heated, and soon burning carbon monoxid appears at the opening of each receiver. The characteristic greenish-blue flame soon appears, showing that the metal is volatilizing. The reduced metal is removed from time to time, the whole operation requiring about eleven hours. The zinc is afterwards redistilled to purify it.

Industrial Applications. Zinc is largely used in "galvanizing" iron, which is done by dipping clean iron into melting zinc. It is also used in the manufacture of brass and other alloys. Zinc dust finds an extensive use in organic chemistry. It is used on a large scale in the manufacture of indigo blue. It is also used as a paint for iron articles.

Compounds. See Exps. 12 and 15 e, Part I.

Uses. The oxid (ZnO) is largely used as a white paint. The chlorid (ZnCl_2) is a strong antiseptic, and is used also in soldering. The sulfate (ZnSO_4) is used in dyeing and calico printing.

CADMIUM

History and Occurrence. Cadmium was discovered by Stromeyer in 1817. It occurs in zinc ores. The first batch of zinc obtained from the zinc smelter mentioned above contains practically all the cadmium; and, since it boils at a lower temperature than zinc, the cadmium distills over first.

Compounds. See Exp. 51, Part I.

ALUMINUM

History and Occurrence. Aluminum was first obtained by Wöhler in 1827, who separated the metal by means of metallic sodium. It is now obtained by an electrolytic method invented by Hall in 1886.

Aluminum occurs almost entirely as a silicate. Clay, mica, slate, etc., are all silicates of aluminum. Its oxid (Al_2O_3) is a common ore, and *cryolite* (AlNa_3F_6) is a very useful mineral. The precious stones, ruby and sapphire, are oxids of aluminum, colored respectively by chromium and cobalt compounds.

Extraction. Hall's method of reducing aluminum compounds is as follows: A large iron receptacle is lined with carbon. Into this extend a number of large carbon electrodes. A mixture consisting of cryolite and fluorite is placed in the receptacle. A strong electric current is then made to pass through the apparatus. The current passes in through the carbon rods, which act as the positive electrode, while the carbon lining acts as the negative electrode. After the flux is melted by the current, aluminum oxid (Al_2O_3) is added at intervals. The melted aluminum collects at the negative electrode, while the oxygen unites with

carbon at the positive. The aluminum, when it has formed in sufficient quantities, is ladled out from the flux.

Applications. On account of its lightness, aluminum is used for the manufacture of utensils where great strength is not required. It forms numerous alloys, one of which, aluminum bronze, is very beautiful.

Compounds of Aluminum. See Exp. 55, Part I.

Manufacture of Aluminum Sulfate. Finely powdered China clay, aluminum silicate, is roasted and then heated with sulfuric acid. In case other clays are used, the iron is precipitated by means of potassium ferrocyanid. The product is sold on the market under the name of "alum cake."

Alum. An alum is a double sulfate of a triad metal and an alkali, and has twelve molecules of water of crystallization. Aluminum sulfate is not readily obtained in crystalline form; hence it is usually crystallized together with potassium or ammonium sulfate, with which it forms crystals of potassium or ammonium aluminum sulfate, $\text{KAl}(\text{SO}_4)_2 \cdot 12\text{H}_2\text{O}$, in the form of regular octahedrons.

The best commercial method for obtaining alum is to roast, together with coal, the bituminous shale of coal beds. This shale is generally rich in aluminum silicate and iron pyrites. The ferric sulfid gives off sulfur, which becomes oxidized to sulfuric acid. Both aluminum and ferric sulfates are formed. This mass is dissolved in water, and evaporated down until the ferric sulfate separates out. Then crude potassium chlorid and potassium sulfate are added, and the mixture is agitated until it is cold. The small crystals then formed are washed with cold water, and recrystallized.

Uses of Alum. Alum is used mostly in dyeing establishments as a mordant. The cloth is placed in an alum solution

to which sodium carbonate has been added until the precipitate first formed has been redissolved. After the cloth is dry the color is applied. The mordant causes the color to stick to the fibers of the material.

Alum is also used in the manufacture of paper.

CHROMIUM

History and Occurrence. Vanquelin and Klaproth discovered, independently of each other, the element chromium in the mineral *crocoisite*. It was named chromium (from $\chi\rho\omega\mu\alpha$, color), because its compounds are all colored.

It is found in crocoisite (PbCrO_4), and in chromite ($\text{FeO}, \text{Cr}_2\text{O}_3$). The green color of emeralds is due to the presence of chromium.

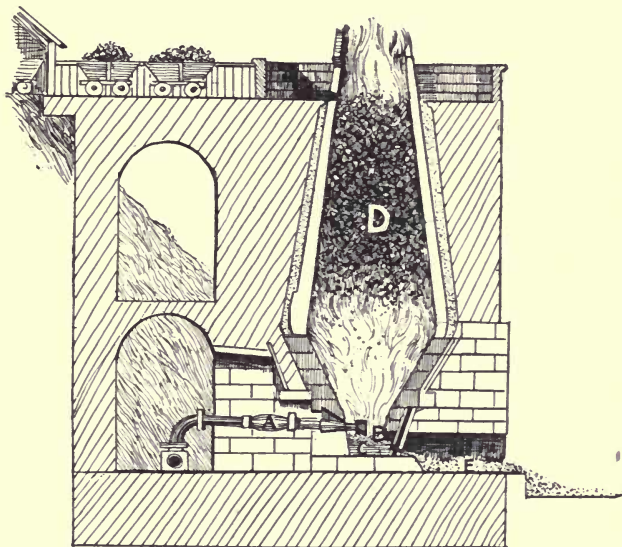
It is a hard, gray, almost infusible metal.

Compounds. Its principal compounds are chrome alum, $\text{CrK}(\text{SO}_4)_2 \cdot 12\text{H}_2\text{O}$, used in dyeing and tanning; lead chromate, PbCrO_4 , a valuable pigment; potassium chromate, K_2CrO_4 ; and potassium bi-chromate, $\text{K}_2\text{Cr}_2\text{O}_7$. The last is the most important commercial product, and is used extensively in bi-chromate batteries and in dyeing.

IRON

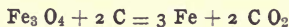
History. It is supposed that iron, the most important and most useful of all the metals, was first extracted from its ores in India. Moses speaks of iron as used by the Hebrews. The Greeks obtained their iron from the Chalybes, a nation living on the shores of the Black Sea. The Romans obtained theirs from Spain and Elba, and from their own mines. Our own iron is mostly obtained from Pennsylvania, although great iron industries have sprung up in Alabama, Tennessee, and Illinois.

Occurrence. Iron is found widely distributed in the earth. It occurs free in certain rocks and in meteorites. The common ores of iron are:—magnetite (Fe_3O_4), haematite (Fe_2O_3), siderite (FeCO_3), and limonite ($2\text{Fe}_2\text{O}_3, 3\text{H}_2\text{O}$).

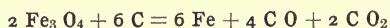


A, Tuyère. B, Hearth. C, Molten iron. D, Limestone, coke, and ore.
E, Sand.

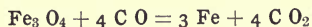
Extraction. It would seem to be a very simple process to obtain iron from its ore. Theoretically, all that is necessary with most ores would be to remove the oxygen. This is done in practice by means of carbon, but difficulties come up that have to be conquered before the final product is obtained. Theoretically, carbon would be the only thing necessary to use to remove the oxygen, thus:—



It is found that carbon monoxid is also formed thus : —



But carbon monoxid is also a reducing agent when the temperature is high enough, and acts upon the ore, thus : —



Iron ore, however, is rarely pure, the impurities being mostly silicates in some form or other. These must be fused or the carbon cannot get a chance to act upon the ore. To effect this, limestone (CaCO_3) is added to the carbon, which is used in the form of coke. The limestone, when heated together with the silicates, produces a flux or molten mass. The process is carried on in a furnace from 75 to 100 ft. high. The ore, coke, and limestone are shoveled into the furnace in layers and ignited. A blast of air is driven through the furnace from the bottom by means of a number of pipes called *tuyères*. The reduction takes place; the iron, from its greater specific gravity, sinks to the bottom of the furnace, and is drawn off into molds of sand three or four feet long and three or four inches wide and deep. This is called “pig iron,” and contains as impurities sulfur, silicon, carbon, manganese, and a small amount of phosphorus.

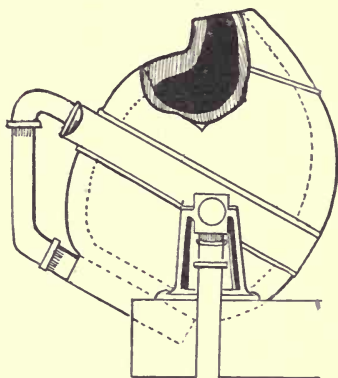
Wrought Iron. The pig iron obtained from the blast furnace is made into wrought iron by removing from it the carbon, of which there is from 2 to 6 per cent. The process is called “puddling.” The pig iron is placed in a reverberatory furnace. The sides of the furnace are lined with a substance containing haematite. At first the iron is melted slowly. Finally it boils, and some of the carbon is removed by the oxygen of the haematite. The heat is then increased, and the molten mass is stirred by workmen by means of iron rods. When the carbon is burned away, the iron becomes pasty, and is removed in large masses on the

ends of the rods. The iron thus obtained is tough. Most of the phosphorus, sulfur, and silicon of the pig iron remains in the furnace in the slag, and is afterwards removed.

Steel. Steel can be made from either pig iron or wrought iron.

To make it from wrought iron, the bars are packed in powdered charcoal, and heated for a number of days at a red heat. The bars are then allowed to cool slowly. The best quality of steel is made in this way. It is called "blister steel," since, when taken from the furnace, the bars are covered with blisters.

The *Bessemer process* is the ordinary way for making steel. Good pig iron is melted in huge vessels, called *converters*, hung on trunnions. A blast of air is blown through the molten iron from the bottom. The flame is watched by an expert until just the right moment, and then the blast is stopped. Carbon is then added to the purified iron by means of molten *spiegeleisen*, a kind of cast iron that contains carbon and manganese. The converter is then emptied, and the steel is cast into ingots.



PROPERTIES OF THE VARIOUS KINDS OF IRON

Pig iron is brittle and cannot be welded. It is the kind of iron used in making castings. *Wrought iron* is tough, malleable, ductile, and can be welded. *Steel* is usually brittle. Its hardness can be changed by tempering; that is, by

heating at various temperatures and suddenly cooling. By allowing hot steel to cool slowly, it can be made almost as soft as wrought iron.

Compounds of Iron. See Exps. 14 *a*, *b*, 15 *f*, and Exp. 56, Part I.

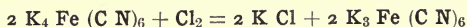
FERROUS SULFATE

One source of ferrous sulfate is the bituminous shale from coal beds. The ferrous sulfate is crystallized out from the solution of the roasted product. (See manufacture of alum.) It is used in the manufacture of black ink, in certain photographic processes, and as a reagent in the laboratory.

POTASSIUM FERROCYANID AND FERRICYANID

When nitrogenous organic matter is fused with potassium hydroxid and iron filings, and the product is treated with water, a solution is obtained which, when purified and evaporated, gives yellow crystals of potassium ferrocyanid, $K_4Fe(CN)_6$. This is commonly called *yellow prussiate of potash*. If treated with ferric chlorid a beautiful blue precipitate, soluble in excess of the chlorid, is formed and is called *Prussian blue*.

If the potassium ferrocyanid is treated with chlorin, the ferricyanid is formed according to the formula:—



When evaporated and crystallized, red crystals of the ferricyanid are obtained. No precipitate with ferric compounds is formed by the ferricyanid, but with ferrous salts a dark blue precipitate called *Turnbull's blue* is formed.

MANGANESE

History and Occurrence. In 1807, Gahn first isolated manganese. Its most abundant and best known ore is pyro-

lusite (MnO_2). The metal itself is reddish-white, oxidizes easily, and must be kept away from contact with air.

Applications. It finds its main use in its compound manganese dioxid (MnO_2), which is a very valuable oxidizing agent. *Spiegeleisen* is an alloy with iron and contains carbon. As has been stated before, spiegeleisen is used in the manufacture of steel from pig iron. Potassium permanganate is also a valuable oxidizing agent. Manganese sulfate is often used as a mordant.

NICKEL AND COBALT

History. Nickel was discovered by Cronstedt in 1751, and cobalt probably by Brandt somewhere about 1735. The two metals both occur in combination with arsenic as arsenids. They are almost always associated, and are separated with difficulty. Of the two metals, nickel is the more important.

Extraction of Nickel. The ore, nickel arsenid, is heated in a reverberatory furnace, thus driving off most of the arsenic. The residue is dissolved in hydrochloric acid; and the impurities, with the exception of cobalt, are removed by various processes. From the remaining solution, the cobalt is precipitated as $\text{Co}(\text{OH})_2$ by bleaching powder. When heated, the cobalt hydroxid becomes cobalt oxid, and is removed by filtering. The nickel solution is then treated with calcium hydrate, and is finally reduced, at a white heat, to nickel by means of carbon. It appears on the market in small cubes.

Uses. Nickel is valuable as an alloy with steel, and with copper and zinc it forms the well-known alloy *German silver*. It is also used for plating iron, for coins, and for laboratory utensils.

Principal Compounds. See Exp. 57, Part I.

TIN

History and Occurrence. Tin was known to the Romans, who obtained it from England after Caesar's conquest. It occurs chiefly as *cassiterite*, or tin stone, from which the tin of commerce is obtained. The oldest tin mines are those of Cornwall. Tin is also found in considerable quantities in Australia and in the Black Hills of South Dakota.

Extraction. The tin ore is first crushed. It is then washed, and the cleansed ore is roasted, to free it from sulfur and arsenic. Mixed with coal and a small quantity of lime, it is then heated. Metallic tin separates out. The metal obtained is somewhat impure; and in order to remove the other metals, it is heated gently. The tin that melts first is ladled out, and this product is further purified by being melted and then stirred with wet wooden sticks. The impurities separate out and the metal is cast into ingots.

Uses. The most important application of tin is in the manufacture of tin plate. Its important alloys are bronze (tin and copper), and plumber's solder (tin and lead). It is also rolled into thin sheets known as tin-foil.

Compounds. See Exp. 54, Part I.

Reducing Action. Stannous chlorid is a very valuable reducing agent in the laboratory.

BISMUTH

History and Occurrence. Bismuth was spoken of by Basil Valentine in the 15th century. Pott first made a careful study of its properties in 1739.

It occurs in nature and often nearly pure. When impure, it is mixed with iron, carbon, and slag, and is then heated in pots. The purer bismuth settles to the bottom. It may

be further purified by melting it on an inclined iron plate, when the pure bismuth melts first and runs off.

Applications. Bismuth is added to certain alloys, to lower their melting point. It is valuable for making fuses that melt at a moderate elevation of temperature, as in automatic fire extinguishers; also for safety fuses and lightning arresters in connection with the use of the electric current.

Compounds. See Exp. 50, Part I.

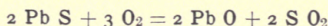
Basic Bismuth Nitrate is extensively used in medicine in the treatment of cholera. It is also used as a cosmetic, and in the manufacture of porcelain.

LEAD

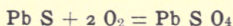
History and Occurrence. Lead is first mentioned in the Book of Job, and Pliny pointed out the distinction between lead and tin. Lead is rarely found free in nature, but occurs in large quantities as *galena* (PbS) in England, Spain, and the United States.

Reduction from the Ore. The ore galena is mixed with lime, and is at first heated gently, a current of air being drawn through the furnace. It is then heated to a higher temperature, and metallic lead is drawn off. The reactions that occur are:—

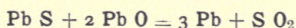
Some lead sulfid is changed to lead oxid,



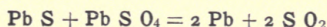
and some to lead sulfate,



Lead oxid and lead sulfate each react upon lead sulfid, forming lead and sulfur dioxid.



and



There is probably always some silver in galena. This, if sufficient in amount, is separated by a process described later (see page 219).

Applications. Lead is used principally for making water pipes, roofing, shot, and various alloys, such as type-metal, solder, and pewter. In making shot, some arsenic is added to the lead. Lead compounds are used in making white lead, and in glass making.

Compounds. Besides the compounds studied in Exp. 53, Part I., there are a few others that should be noted. These are red lead (Pb_3O_4); white lead, which is a basic lead carbonate; and lead acetate, commonly called "sugar of lead."

Red Lead is made by heating litharge (PbO) on trays in a reverberatory furnace.

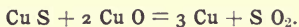
White Lead is made by placing rolls of perforated sheet lead in earthenware pots partly filled with diluted vinegar (acetic acid). The lead is not in contact with the vinegar. The pots are covered with tan bark or manure. In time the acid fumes change the lead to lead acetate, and the carbon dioxide from the tan bark or manure changes this to the carbonate.

Lead Acetate is made by dissolving litharge in acetic acid.

COPPER

History and Occurrence. Copper was probably the first metal used by mankind, having been used by pre-historic man in making his weapons. It is very abundant, which is fortunate, since it is one of the most useful metals we have. It occurs native in the Lake Superior regions, the mines there furnishing almost chemically pure copper. In Montana, Idaho, and Arizona, vast quantities of copper occur, associated with the precious metals. It is also found as the oxide in Russia and Australia, and as the sulfide in England.

Extraction. When the oxid ore is heated with carbon, the metal separates out. The sulfid ore, however, must first be roasted in order that part of the sulfid may be changed to oxid. Then the mixture of sulfid and oxid is heated further in a reverberatory furnace, where the following reaction occurs.



Industrial Applications. Since the development of electricity, copper has become one of the staple products, on account of its use as an electric conductor. Its use in the preparation of the alloys, brass and bronze, has already been mentioned. As sheathing for ships, etc., large quantities are used. Copper coins are familiar to us all.

Compounds. See Exps. 4; 9 g; 39; Part I.

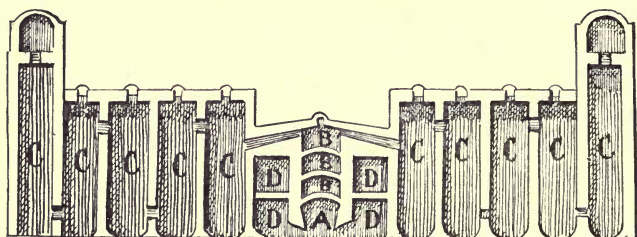
Copper sulfate is an important commercial product. The sulfid is gently roasted, and the copper sulfate thus formed is dissolved out with water. It is also formed as a by-product in the purifying of gold and silver. It is used in the preparation of copper arsenite (Paris green), in the manufacture of other pigments, and also in calico printing.

MERCURY

History. Mercury is mentioned by the later Greek writers under the name ὑδράργυρος (from ὕδωρ, water, and ἄργυρος, silver). Pliny called it *hydrargyrum*. The alchemists believed that mercury was a component of all metals. Braune was the first to recognize it as a true metal. In the winter of 1759, he solidified it by using snow and nitric acid to produce the necessary cold.

Occurrence and Extraction. Mercury occurs free in small quantities but is found chiefly in combination with sulfur as *cinnabar* (HgS), in Idria, Carroliia, in Spain, and in California.

The ore mixed with lime is placed in a furnace into which a current of air may pass. When the ore is heated, the sulfur becomes sulfur dioxid; this, together with the vapor of mercury, passes into a series of chambers. The mercury vapor is here condensed, while the sulfur dioxid passes



A, Fire box. B, Ore. C, Condensing chambers. D, Chambers through which air is admitted.

on. In the last chamber falling water condenses the last portion of the vapor. The mercury thus obtained is partially purified by being filtered through linen, and then by the action of dilute nitric acid. It is then further purified by redistillation, and by being forced through chamois skin. The metal is usually stored in wrought iron bottles.

Compounds. See Exp. 52, Part I.

Uses. In mining, mercury is used in extracting gold and silver from their native mixtures. In the arts, it is used in making barometers, thermometers, mirrors, etc. Its compounds are also used in the arts and in medicine. The chemist and physicist would be at a great disadvantage without this convenient metal.

Amalgams. Mercury forms amalgams with all well-known metals except iron. The most important are with potassium, sodium, cadmium, copper, silver, and gold. It also forms an amalgam with the metalloid ammonium (see Exp. 42, Part I.).

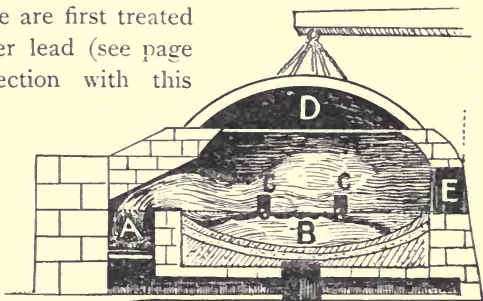
SILVER

History and Occurrence. Silver was one of the seven metals of the ancients. It occurs in nature, masses of it having been found weighing as much as 800 lbs. In a Norwegian museum is a mass weighing 500 lbs. Its most important ores are *argentite* (Ag_2S), *pyrargentite* (Ag_3SbS_3) and *horn silver* (AgCl). Galena generally contains small quantities of silver sulfid. The best known silver mines are in the western part of the United States, in Mexico, in South America, and in Australia.

Extraction. The principal methods of obtaining silver from its ores are the *amalgamation* process and *cupellation* process.

In the *amalgamation* process, the ore is first roasted with common salt, giving silver chlorid. This is then agitated with scrap iron and mercury, giving ferrous chlorid and an amalgam. The amalgam is washed and filtered through canvas, to remove the excess of mercury. The pasty amalgam that is left is then distilled, leaving the pure silver behind.

The *cupellation* process is adapted to lead ores containing silver. These are first treated as described under lead (see page 216). In connection with this process, the following preliminary treatment is necessary. The lead obtained is melted and allowed to cool. The very pure



A, Fire box. B, Hearth. C, Tuyères. D, Hood.
E, Opening for introducing materials.

lead crystallizes at the top, and is skimmed off. New alloy is added, melted, and the same operation is continued. This is repeated until fairly pure silver remains. The silver thus obtained is then placed in a shallow furnace, called a *cupel*, made of a mixture of clay and lime. The alloy is subjected to a high temperature, while a current of air is driven over the molten mass. This oxidizes the lead to litharge, and pure silver remains.

Applications. Silver has been used for coinage from earliest times. This, and the manufacture of silver tableware, ornaments, etc., consume most of the silver of the world. A large amount of silver salts is used in photography.

Photography. The fact that light blackens many silver compounds has led to the foundation and development of the art of photography. The first "light pictures" were made by Wedgewood in 1802. These were prints of leaves on paper moistened with silver nitrate. Exposure to light, however, soon blackened and destroyed these prints. In 1826, Daguerre found a process by which the image could be preserved. Our old family Daguerreotypes are the surviving relics of this process.

The present dry plate method makes use of two processes: first, the making of the negative upon a sensitive film; second, the transferring of the picture to a second sensitive film on paper, called the positive.

The negative is a film usually composed of gelatine and silver bromid (AgBr) with traces of other silver salts, adhering to a plate of glass or some other transparent substance. When placed in the camera and exposed to light, no picture is shown until the plate is "developed." This is done by various methods. One good developing solution consists of pyrogallic acid and sodium carbonate. After

the picture has been brought out by the "developer," the unchanged silver bromid must be removed. This is done by means of a solution of sodium thiosulfate ($\text{Na}_2\text{S}_2\text{O}_3$). The products of the chemical changes are washed away with water, and there is left only silver on the plate. The varying amount of silver on different parts of the plate produces the dark and light portions.

The picture must now be transferred to paper. One kind of sensitive paper has on one side of it a film of albumen and silver chlorid. The negative is placed over the paper, thus leaving the sensitive silver chlorid unprotected from the light where the silver is least deposited on the negative. By allowing this to stand in sunlight, the picture is "printed." It must then be "toned" and "fixed." In "toning," part of the silver chlorid is usually replaced by gold from a solution of auric chlorid (AuCl_3), or by platinum from a salt of that metal. The unchanged silver chlorid is removed by sodium thiosulfate solution, just as was done in the negative. The picture is then washed and mounted.

GOLD

History and Occurrence. Gold was considered by the alchemists to be the most perfect metal, and many spent their lives vainly trying to transmute the baser metals into it.

Gold is usually found native. The most famous gold fields have been those of California, Australia, South Africa, and Alaska. It usually occurs in alluvial deposits, or else in quartz veins.

Extraction. The auriferous gravel or clay is washed; the lighter materials are thus allowed to be carried away, while the heavy gold remains behind. Gold-bearing quartz is crushed, then mixed with mercury, and amalgam of gold is

formed. The mercury is distilled off, leaving the gold. Native gold usually contains silver. After both metals are dissolved in aqua regia (see Exp. 61 c, Part I.) the silver is precipitated by hydrochloric acid. The auric chlorid is then treated with ferrous sulfate, giving metallic gold, which is collected and melted into ingots. Both gold and silver are frequently reduced in blast furnaces in connection with ores of lead, the precious metals being concentrated in the lead ingot formed. When gold and silver are mainly associated with copper, they are obtained in the copper *matte*, which is produced by a somewhat complex process of reduction in the reverberatory furnace.

Uses. Coinage, jewelry, and plate consume most of the gold supply. It is also used in making gold leaf, and in making ruby glass.

Auric chlorid is obtained by dissolving gold in aqua regia, and evaporating carefully. The deliquescent salt that is formed is used extensively in photography.

Other Gold Compounds are the oxid (Au_2O_3), the sulfid (Au_2S_3), and a remarkably beautiful compound called the Purple of Cassius, which is probably a double stannate of gold and tin.

PLATINUM

History. Platinum was first noticed in the sixteenth century by Scaliger, but not until two hundred years later did it become well known. It is found principally in the Ural mountains, California, Mexico, and Australia.

Extraction. Platinum is extracted from its ores by being treated with aqua regia. This dissolves the platinum together with certain rare metals. After the solution has been evaporated to dryness, the rarer chlorids are decomposed by heat. The platinum chlorid is then dissolved in water, pre-

cipitated with ammonium chlorid, and heated. A spongy mass of metallic platinum results. This is then fused to a button by the oxy-hydrogen blowpipe.

Uses. Platinum is especially valuable for crucibles, dishes, and other utensils of the chemical laboratory. It also has numerous uses in electricity.

QUALITATIVE ANALYSIS

QUALITATIVE ANALYSIS

THE term *qualitative analysis* signifies the art of finding out what elements are present in any given chemical compound or mixture. Qualitative analysis is not concerned with the proportions in which these elements may be present in the compound. Such determinations belong to *quantitative analysis*. Qualitative analysis has for its basis the fact that the elements fall into natural groups when considered with respect to the compounds they form. Thus lead, silver, and mercurous salts are naturally grouped together, because, from solutions of their soluble salts, hydrochloric acid precipitates them as chlorids. Similarly we may group together the elements that are precipitated by other reagents. By experiment we may find out what precipitate is obtained by any given reagent from any given solution of a salt of a single metal. Similarly, we may find solvents that will separate precipitates which have been formed by a reagent acting upon a solution containing more than one metal. In like manner we may discover the precipitates formed by reagents with the salts of any acid. Besides precipitates, we may use other tests, such as flame colorations, to distinguish the elements or acid radicals from one another.

The following table gives, with their solubilities, the important precipitates of the various elements.

	NaOH	NH ₄ OH	Na ₂ CO ₃	(NH ₄) ₂ CO ₃	HCl	H ₂ SO ₄	H ₂ S	(NH ₄) ₂ S	Na ₂ HPO ₄	KCN	K ₂ CrO ₄	Special tests
Lead	W. sol. ex.	W. insol. ex.	W.	W.	W. sol. H ₂ O	W.	Blk. sol. dil. HNO ₃	Blk. sol. dil. HNO ₃		W. sol. HNO ₃	Y. sol. NaOH	KI gives Y.
Silver	Br. sol. NH ₄ OH	W. in neut. sol. NH ₄ OH			W. sol. NH ₄ OH reprecip. HNO ₃		Blk. sol. HNO ₃	Blk. sol. HNO ₃	Y. sol. NH ₄ OH	W. sol. ex.	R. sol. NH ₄ OH	Copper gives metallic silver
Mercurous	Blk.	Blk.			W. turns blk. with NH ₄ OH	Gray	Blk. insol. acids	Blk. insol. acids			R. insol. NH ₄ OH	KI gives R.
Mercuric	Y. sol. warm acids		Br.				1st W. then R. " Blk.					
Copper	Bl. turns blk. on boiling	Gr. Bl. sol. ex.	Gr. sol. KCN	Gr. Bl. sol. ex.			Blk. sol. warm dil. HNO ₃	Blk. sol. warm dil. HNO ₃				K ₄ Fe(CN) ₆ gives Br.
Bismuth	W. insol. ex. sol. HNO ₃	W.	W.	W.			Br. sol. warm dil. HNO ₃	Br. sol. warm dil. HNO ₃			Y. sol. dil. HNO ₃	Water gives W.
Cadmium	W. insol. ex.	W. only in conc. sol. ex.	W. insol. ex.	W. insol. ex.			Y. sol. warm dil. HNO ₃	Y. sol. warm dil. HNO ₃		W. sol. ex.	Y. sol. HNO ₃	

	NaOH	NH ₄ OH	Na ₂ CO ₃	(NH ₄) ₂ CO ₃	HCl	H ₂ SO ₄	H ₂ S	(NH ₄) ₂ S	Na ₂ HPO ₄	KCN	K ₂ CrO ₄	Special tests
Cobalt	Bl. turns R.	Bl. sol. ex.	R. insol. ex.	R. sol. ex.				Blk. insol. ex.				KNO ₃ gives Y.
Manganese	W. turns Br.		W. turns Br.	W. turns Br.				P. sol. dil. acids				
Zinc	W. sol. ex.	W. sol. ex.	W.	W. sol. ex.				W. sol. dil. acids				
Barium	W.	W.	W. sol. HCl.	W.		W. insol. all dil. acids.					Y. sol. HCl, HNO ₃	
Calcium	W.	W.	W. sol. dil. acids	W. sol. dil. acids		W. sol. in hot (NH ₄) ₂ SO ₄			W.			Ammonic Oxalate gives W.
Strontium	W.		W. sol. HCl, HNO ₃	W. sol. HCl, HNO ₃		W.					Y. in neut. sols. No precip. in acid sols.	
Magnesium	W. in absence ammoniac salts.	W. in absence ammoniac salts.	W. in absence ammoniac salts.	W. in absence ammoniac salts.					W. crystal-line			
Sodium	Gives yellow color to the flame.											
Potassium	Gives violet color to the flame.											
Ammonium	Gives off ammonia when heated with alkalis.											

W, white; Y, yellow; R, red; Blk, black; Bl, blue; Gr, green; O, orange; Br, brown; P, pink; sol, soluble; insol, insoluble; dil, dilute; conc, concentrated; neut, neutral.

By consulting the foregoing tables, it is possible to arrange the compounds of the metals in groups according to their similarities, thus : —

GROUP I.

Lead	}	Precipitated by hydrochloric acid.
Silver		
Mercurous		

GROUP II.

Mercuric	}	A	}	Precipitated by hydrogen sulfid.
Copper				
Bismuth				
Cadmium				
Antimony	}	B		
Arsenic				
Tin				

GROUP III.

Iron	}	Precipitated by ammonium hydroxid.
Aluminum		
Chromium		

GROUP IV.

Nickel	}	Precipitated by ammonium sulfid.
Cobalt		
Manganese		
Zinc		

GROUP V.

Barium	}	Precipitated by ammonium carbonate.
Calcium		
Strontium		

GROUP VI.

Magnesium. Precipitated by acid sodium phosphate.

GROUP VII.

Sodium	}	Recognized by flame colorations.
Potassium		
Lithium		

Ammonium. Recognized by special tests.

How to Analyze an Inorganic Substance

EXAMINATION FOR BASES

If the substance is not in solution, it may be dissolved in water, in acids, or, after fusing with sodium carbonate, it may be dissolved in acids.

To the liquid containing the mixture to be analyzed, add dilute *hydrochloric acid*. Allow the precipitate, if one is formed, to settle; then add a drop or so of the acid, to make sure that no further precipitate can be formed. Filter, wash with cold water, and *save* the filtrate. The precipitate contains Group I.

GROUP I.

Break a hole in the filter, and wash the precipitate into a test tube by means of a fine stream of water from a wash bottle. Boil and filter again. The filtrate from this contains the soluble *lead chlorid*, if any is present. This may be tested for **lead** by *potassium chromate*, which gives a yellow precipitate soluble in *sodium hydroxid*.

To the precipitate (if there is any) remaining on the filter, add *ammonium hydroxid*, and collect the resulting filtrate in a beaker. If the precipitate turns black, a **mercurous** compound is present. To the filtrate add *nitric acid*. If any *silver chlorid* was dissolved, it will be reprecipitated by the nitric acid, thus indicating the presence of **silver**.

The filtrate from Group I. may contain the compounds of all elements belonging to the succeeding groups. Dilute this with water, and heat to boiling; then, under a hood, pass a stream of *hydrogen sulfid* through the solution until it is saturated. The solution is saturated, if, on removing the beaker and blowing away the gas from above the liquid, it

still smells of hydrogen sulfid after a lapse of one or two minutes. Then filter, and *save* the filtrate. Wash the precipitate until the wash water is no longer acid. The precipitate contains Group II.

GROUP II.

Remove the precipitate to a porcelain dish, add enough concentrated *ammonia* to cover the mixed sulfids, and warm, stirring the mixture continually. Then add a little *yellow ammonium sulfid* together with the *ordinary kind*, until the mixture smells distinctly of the sulfid. After warming a few minutes, allow the mixture to stand about fifteen minutes. Then filter, and wash the precipitate with hot water. The precipitate may contain any or all of A. of Group II., i.e., *mercuric, copper, bismuth, or cadmium sulfids*. The filtrate may contain any or all of B. of Group II., i.e., *antimony, arsenic, or tin sulfids*.

A. OF GROUP II.

Remove the precipitate to a porcelain dish, and add enough *dilute nitric acid* to cover the mixed sulfids about twice over. Boil, and stir the mixture continually until all solvent action has ceased. Filter and wash. The precipitate contains *mercuric sulfid* and any *lead* that may have been dissolved by the water from Group I. The presence of the mercuric sulfid is confirmed by adding a few drops of *aqua regia* to the precipitate, diluting, and then adding a little *stannous chlorid*. If a white precipitate, which blackens on the addition of *ammonia*, is formed, the indication is that **mercury** is present.

The filtrate may contain *copper, bismuth, and cadmium nitrates*. Add *ammonium hydroxid* to excess. A white flocculent precipitate indicates **bismuth**, and a blue color in the

solution indicates **copper**. Filter, and add *potassium cyanid* until the blue color disappears; then pass a current of *hydrogen sulfid* through the solution. A yellow precipitate indicates **cadmium**.

B. OF GROUP II.

To the filtrate obtained from the mixed sulfids in A., add *dilute hydrochloric acid*. If the precipitate is milky, probably nothing but sulfur, formed by the decomposition of the alkaline sulfid, is present. If the precipitate is lemon yellow, or orange, at least one sulfid of the metals of B. is present. Filter, remove the precipitate to a porcelain dish, and boil it with concentrated *hydrochloric acid*. Hydrochloric acid decomposes the *sulfids of antimony and tin*, leaving the *arsenic sulfid*. Filter. The precipitate remaining may be confirmed as *arsenic* by removing it to a porcelain dish, and adding concentrated *hydrochloric acid* and a few crystals of *potassium chlorate*. Evaporate off the excess of acid; then add *ammonium chlorid*, an excess of *ammonium hydroxid*, and *magnesium sulfate*. If **arsenic** is present, needle-shaped crystals of *ammonium magnesium arsenate* will be deposited.

Dilute the filtrate with water, and add a strip of zinc and a strip of platinum. A black coating upon the platinum indicates **antimony**. If **tin** is present, a gray powder, partly on the bottom of the beaker and partly on the zinc, is formed. This powder may be removed and dissolved in concentrated *hydrochloric acid*. *Mercuric chlorid*, added to the solution, gives a white precipitate of *mercurous chlorid*. This precipitate soon becomes grayish, owing to its reduction to metallic mercury.

The filtrate from Group II. may contain compounds of elements belonging to the succeeding groups. Boil the filtrate until all traces of hydrogen sulfid have been removed.

Then add a few drops of *nitric acid*, to change any ferrous compounds present to ferric. To the hot solution, add a moderate quantity of *ammonium chlorid*. While stirring the solution, add *ammonium hydroxid** in small quantities until a distinct odor of ammonia is perceptible. (The ammonium chlorid is added, to keep in solution the compounds of metals belonging to succeeding groups.) Filter and wash. *Save* the filtrate. The precipitate contains Group III.

GROUP III.

If the precipitate is white, only *aluminum* is present; if it is grayish-green or blue, *chromium* is present; if it is brown, *iron* is present. Dissolve the precipitate, on the filter, by adding *dilute hydrochloric acid* drop by drop. Dilute the resulting solution with water, add *sodium hydroxid* to excess, and boil. The chromium and iron are thus precipitated, while the aluminum hydroxid formed is dissolved by the alkali. When the precipitate has settled, filter. The precipitate may be examined for chromium by fusing a small quantity with a mixture of *potassium nitrate* and *sodium carbonate* upon a piece of platinum foil, dissolving the mass thus obtained in water, adding a little *acetic acid* to liberate the carbon dioxid, and then adding a little *lead acetate*. A yellow precipitate indicates **chromium**. **Iron** in the precipitate from sodium hydroxid is shown by dissolving a little of it in *hydrochloric acid* and testing it with *potassium ferrocyanid*.

The filtrate containing the excess of sodium hydroxid is then tested for aluminum by acidifying with *hydrochloric acid*, and adding *ammonium hydroxid*. A white flocculent precipitate indicates **aluminum**.

* It is assumed that there are no phosphates or oxalates present in the solution. Their presence would make necessary a different treatment of the solution.

The filtrate obtained from Group III. may contain compounds of elements belonging to succeeding groups. Transfer the solution to a flask, and heat it to boiling. Add *ammonium sulfid* to the hot solution, drop by drop, as long as any precipitate is formed, shaking the flask continually. Allow the precipitate to settle, filter, and wash. *Save* the filtrate. The precipitate may contain any or all of the *sulfids of nickel, cobalt, manganese, and zinc*, i.e., Group IV.

GROUP IV.

Place the precipitate, paper and all, in a porcelain dish, and add *cold dilute hydrochloric acid*. This decomposes the sulfids of *manganese* and *zinc*, and leaves the sulfids of *nickel* and *cobalt* unchanged. Filter and wash. Treat the precipitate with a little *aqua regia*, and evaporate it to dryness. Dissolve the resulting chlorids in water, and add a medium sized piece of *potassium nitrite*. Yellow crystals of *potassium cobalt nitrite* will form, if **cobalt** is present. Filter these off, and test the filtrate for **nickel** by adding *sodium hydroxid*.

The filtrate, obtained by treating the original mixed sulfids with dilute hydrochloric acid, is now boiled until all hydrogen sulfid is expelled. *Sodium hydroxid* is then added, and, if **manganese** is present, a white gelatinous precipitate of manganese hydroxid is formed. The precipitate rapidly turns brown. The filtrate from this contains the zinc hydroxid, if any is present, dissolved in excess of sodium hydroxid. Pass *hydrogen sulfid* through this. A white precipitate, soluble in hydrochloric acid, indicates **zinc**.

The filtrate obtained from the precipitate of Group IV. may contain any or all of the compounds of metals belonging to the succeeding groups. Boil the filtrate until all ammonium sulfid is decomposed, and filter. Then add a little *ammonium chlorid*, after which add *ammonium carbonate*

as long as any precipitate is formed; then warm gently. The precipitate may contain any or all of the carbonates of Group V. Filter and wash. *Save* the filtrate.

GROUP V.

Wash the precipitate into the apex of the filter, and then add *acetic acid*, catching in a test tube the solution that runs through. To this add *potassium chromate*; barium chromate is formed, if **barium** is present. Filter, and to the filtrate add dilute *ammonium hydroxid*. Then add *ammonium carbonate* to reprecipitate the calcium and strontium, and filter again. Wash until the precipitate is white, and then redissolve it in *acetic acid*. Add to this a dilute solution of *ammonium sulfate* (1 part sulfate to 12 parts of water), and allow the precipitate, if one is formed, to settle. A white precipitate indicates **strontium**. Then filter, and add a little *acetic acid* and *ammonium oxalate*. A white precipitate, soluble in *hydrochloric acid* and reprecipitated by *ammonium hydroxid*, indicates **calcium**.

The filtrate obtained from the precipitation of Group V. may contain the compounds of metals belonging to Groups VI. and VII. The reagent for Group VI. is *acid sodium phosphate*. Since this reagent contains sodium, it would evidently be useless to test for sodium in the original solution, after once having added the reagent. It is hence necessary to divide the filtrate from Group V. into two parts, and examine one part for Group VI., and the other for Group VII.

GROUP VI.

Place one portion of the filtrate from Group V. in a large test tube, and add *ammonium hydroxid* and *acid sodium phosphate*. Shake the solution well, scratch the inside of

the tube with a glass rod, and allow the mixture to stand. A white crystalline precipitate indicates **magnesium**.

GROUP VII.

The remaining part of the filtrate from Group V. is evaporated to dryness, and is then heated so that all ammonium compounds are decomposed. It may then be tested for potassium and sodium by the color that the residue gives to a non-luminous flame. If **sodium** is present, the characteristic sodium flame will be seen. This flame, however, will obscure any coloration due to potassium; therefore it is necessary to view the flame through a piece of cobalt glass. The cobalt glass shuts off the yellow sodium rays, and allows the violet **potassium** rays to pass through. **Lithium** gives a beautiful red color to the flame. In order to test for the presence of ammonium compounds, it is necessary to use the original mixture. To a small test of this, add *sodium hydroxid*. If, on boiling, the steam turns red litmus paper blue, **ammonium** is present.

EXAMINATION FOR ACIDS

In examining a solution for acid radicals, we do not have the advantage of such a well-defined and systematic scheme as we have in examining it for bases. We are, in this case, obliged to depend upon a knowledge of the relations that various acids bear to various bases, and also upon numerous special tests. For instance, if we find a lead, silver, or mercurous salt in a solution, we know that no halogen can be present; for the halogen would then precipitate the metal. Likewise, no sulfate can be found in a solution containing a salt of barium.

The special tests for the various common acids may be found by referring to the following table.

Salts of	AgNO_3	Pb Ac_2	$\text{Hg}_2(\text{NO}_3)_2$	BaCl_2	
HCl	W. sol. NH_4OH	W.	W. turns Blk. with NH_4OH		
HBr	Yellowish sol. KCN	W. sol. dil. HNO_3	Y. W.		
HI	Y. sol. KCN	Y. sol. dil. HNO_3	Gr. Y. sol. KI		$\text{Hg}(\text{NO}_3)_2$ gives R. sol. KI
H_2SO_3	W.	W. sol. dil. HNO_3		W. sol. dil. HCl	
H_2SO_4		W. sol. tartaric acid and NH_4OH		W. insol. all dil. acids	
H_3PO_4	Y. sol. NH_4OH , HNO_3	W. sol. HNO_3		W. sol. HCl, HNO_3	
HCN	W. sol. NH_4OH	Add small amount FeCl_3 and FeSO_4 ; make alkaline with NH_4OH ; warm and acidify with HCl. Turns blue.			
HF	Gives HF with H_2SO_4 .				
H_2S	Various precipitates with metals.				
HNO_3	Two or three cc. conc. H_2SO_4 , poured carefully to bottom of sol. in test tube and FeSO_4 sol. added, gives brown ring.				
H_3BO_3	Colors alcohol flame green.				
H_2CO_3	Gives CO_2 with acids.				
H_4SiO_4	Give insoluble skeleton in microcosmic bead.				

APPENDIX I.

Metric System

LINEAR MEASURE

10 millimeters	=	1 centimeter
10 centimeters	=	1 decimeter
10 decimeters	=	1 meter
10 meters	=	1 dekameter
10 dekameters	=	1 hektometer
10 hektometers	=	1 kilometer

CAPACITY

100 centiliters	=	1 liter
100 liters	=	1 hektoliter
1 liter	=	1 cubic decimeter
	=	1000 cubic centimeters

WEIGHT

10 milligrams	=	1 centigram
10 centigrams	=	1 decigram
10 decigrams	=	1 gram
10 grams	=	1 dekagram
10 dekagrams	=	1 hektogram
10 hektograms	=	1 kilogram
1000 milligrams	=	1 gram
1000 grams	=	1 kilogram
1000 kilograms	=	1 tonneau

1 cubic centimeter of water at 4° C. weighs 1 gram.

1 cubic decimeter of water at 4° C. weighs 1 kilogram.

APPENDIX II.

Manipulations

1. *To Cut Glass Tubes.* Mark the glass with a file at the required point, then, seizing the tube with the hands so that the thumbs are opposite the mark, break by pulling with the hands and at the same time pressing outward with the thumbs.

To cut large glass tubes, make a mark with a file as before. Take two pieces of filter paper long enough to go around the tube several times, and fold them twice so as to form a strip of four thicknesses, 3 cm. wide. After wetting the papers thoroughly, wrap them around the tube about 4 mm. from the mark, one on each side of it. Be careful to have the edges of the wet filter paper straight and touching the glass at all points. Then rotate the wrapped portion of the tube in the hot flame of a Bunsen burner, when, after a few moments, the tube will break off evenly.

2. *To Bend Glass Tubes.* A fish-tail gas flame should be used. Hold the tube lengthwise in the flame just above the dark part, at the same time rotating it with the fingers. When the glass has softened enough, *remove it from the flame and quickly bend it to the desired angle.* Allow the tube to cool with the soot upon it.

3. *To Insert a Glass Tube into a Rubber Stopper.* Wet both glass and hole. The tube can then be easily pushed in the required distance. Be very careful in pushing tubes into rubber stoppers, since a glass tube broken while held in the hand in this position sometimes makes an ugly wound.

4. *To Bore a Hole through Glass.* Procure a rat-tail file about 4 mm. in diameter. Break the file and grind the

broken end down on two sides to a convex edge. Moisten the file and the glass with the following solution:—

1 oz. Camphor Gum,

1½ oz. Oil of Turpentine.

Directly under the point where the hole is desired, place a small piece of hard wood about 1 sq. cm. on the base with an upper surface of about 4 sq. mm. Apply the end of the file, using it as you would an awl. When the glass is pierced, grind the hole out to the required size by means of a smaller rat-tail file. Moisten the file occasionally with the solution.

5. *To Dry a Flask.* First rinse the flask with a little alcohol. Take a tube several inches longer than the flask, and insert it to the bottom. Then force air through the tube with a bellows, at the same time rotating the tube in the flame of a Bunsen burner. The alcohol will be vaporized, and the flask will be dry in a few moments.

6. *To Make a Wash Bottle.* Fit a two-holed rubber stopper with two glass tubes. Let one reach to the bottom of a liter flask, and have its outer end bent at an angle of 45°. The outer arm of this tube should be about 3 cm. long, and should be connected by means of rubber tubing with a short glass tube of the same bore drawn out to a diameter of about 1 mm. The second tube should extend only through the stopper, and its outer portion should be bent at an angle of 135°. Round off the ends of both tubes in the Bunsen flame. To make the flask easy to handle when containing hot water, bind a strip of cork around its neck.

7. *Filters.* Filter papers are usually folded, first along one diameter and then along a line perpendicular to this. When the filter is opened, one side has three thicknesses of paper, and the other side has one.

8. *To Use a Blowpipe.* The use of the blowpipe requires a constant passage of a stream of air through it. To pro-

duce this, fill the mouth with air, at the same time breathing through the nose. The muscular action of the cheeks will force a constant current of air through the blowpipe. Do not attempt to breathe through the mouth, but replenish the air in the mouth from the lungs. A little practice will make you so expert that you will be able to blow a constant flame for many minutes.

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